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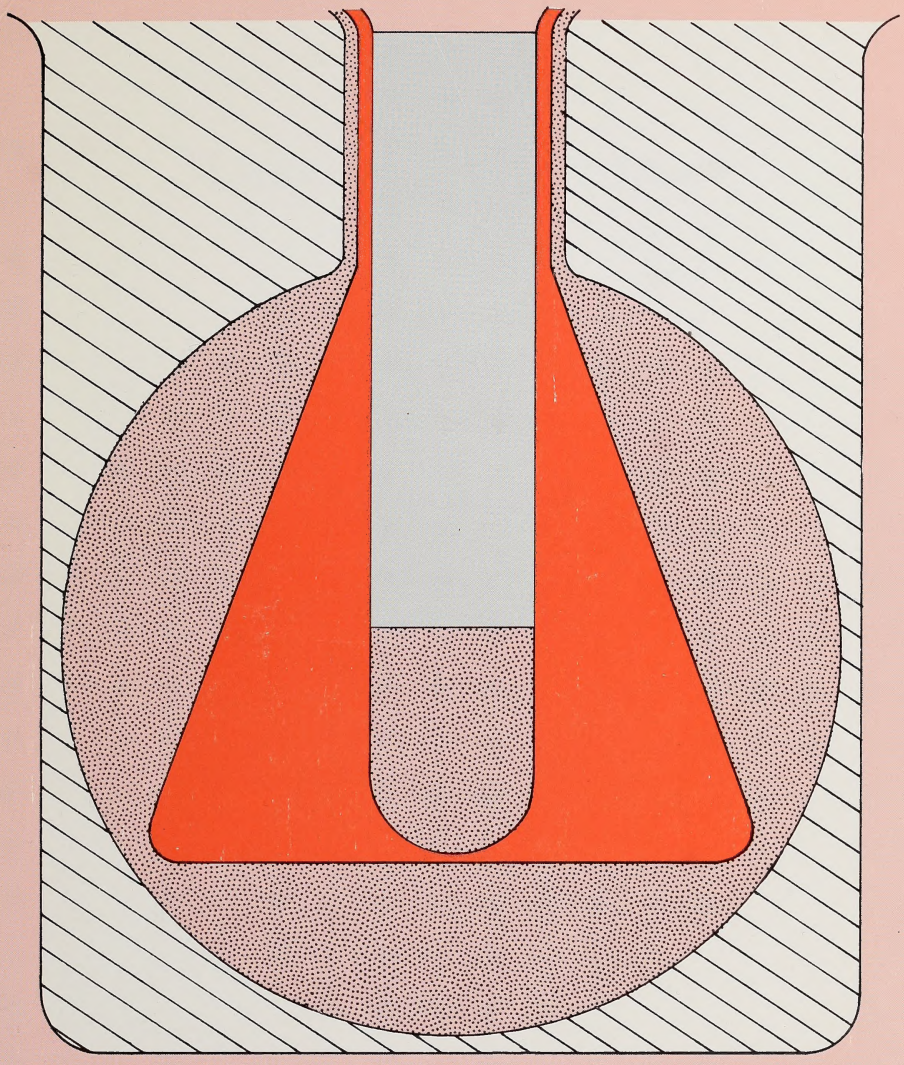
# CHEMISTRY



# UPGRADING

Alberta
EDUCATION

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# Chemistry Upgrading

## Lessons A-J

Distance Learning

Alberta EDUCATION

Chemistry Upgrading
Student Module
Lessons A-J and Answers
Alberta Correspondence School
ISBN No. 0-7741-0069-9

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# INTRODUCTION

## Purpose and Outline of the Course

The Chemistry 30A course has been designed to provide the adult student with a grounding in the concepts of modern chemistry. Many adult students who have been away from the classroom for several years do not have sufficient background in the modern chemical principles to enable them to attempt the Chemistry 30 course. It is hoped that the Chemistry Upgrading lessons will help these students to prepare adequately for the Chemistry 30 final examination.

## Reference Material

All material in this course is self contained, so no textbook is required.

The book which will be later used for the Chemistry 30 part also has many sections dealing with topics covered in lessons A to J. This textbook is:

Chemistry by O'Connor et. al.

**General Instructions**

1. As the first ten lessons do not constitute a credit course and as there will be no test to promote the students into the Chemistry 30 material, you will be expected to rely mainly upon your own endeavor and self-direction to obtain the maximum benefit from the course. It is reasonable to expect that the course will take you a longer time than most correspondence courses as the single Chemistry upgrading course encompasses the amount of chemistry that is normally covered in two years at school.

2. Problems constitute a major part of chemistry. Orderly procedure is absolutely necessary when solving problems. The following steps are recommended.

   (a) Master the typical illustrative examples given in the lesson.

   (b) Read the assigned problem through at least twice, noting the information **given** and the information which is **required.**

   (c) Decide which relationship between the known and required (e.g. the chemical law that links the two) must be used in order to solve the problem. If the relationship can be expressed in the formula form, write down the formula.

   (d) Substitute the known values from the problem, making sure that all values are in **consistent units.**

   You can often make your mathematical work easier by cancellation or some other short cuts.

   (e) Solve the problem and make sure your answer is expressed in the appropriate units.

   (f) Ask yourself if the answer is reasonable in view of the information supplied in the problem and if it answers the question posed by the problem.

NOTE: Keep this list in front of you when initially solving problems. In time the above steps will occur to you automatically when you are faced with a problem.

As mentioned before, the course is self-contained, but if you wish to do extra reading, the textbook by O'Connor will supplement some areas of some of the lessons. It is not common that one chapter in the textbook would completely cover one lesson, however, if you make good use of the Table of Contents and the Index, you should find many sections of interesting reading.

## Outline of Chemistry Upgrading Course (A-J)

Lesson A

A. Systems of Measurements
B. The Concept of Matter and Energy
C. Density
D. Scientific Notation
E. Significant Figures

Lesson B: Classification of Matter

A. Physical and Chemical Properties
B. Physical and Chemical Changes
C. Classification of Matter
    1. Homogeneous or Heterogeneous
    2. Metals or Nonmetals
    3. Mixtures or Pure Substances
    4. Ionic or Molecular Substances
D. Mixtures, Pure Substances, Elements and Compounds
E. Energy and Chemical Changes
F. Symbols
G. Ways of Separating Matter

Lesson C

A. Experiment 1 - Black Box Experiment
B. Development of an Atomic Model
C. Rutherford's Experiments
D. Periodic Law
E. Atomic Theory
F. Gay Lussac's Law of Volumes
G. Avogadro's Principle

Lesson D

A. Subscripts and Coefficients
B. Writing Formulas
C. Naming of Compounds
D. Molecular and Atomic Masses

Lesson E: Chemical Equations

A. Writing Chemical Equations
   1. Qualitative
   2. Quantitative

B. Balancing of Chemical Equations
   1. Using Moles of Atoms
   2. Using Atoms
   3. Conservation of Mass

C. Using Equations in Problem Solving

D. Types of Chemical Reactions
   1. Synthesis Reaction
   2. Decomposition Reaction
   3. Rearrangement or Replacement Reaction
   4. Metal-ion Reaction
   5. Ion-ion Reaction

Lesson F: Chemical Bonding

A. Intra-molecular Bonding
   1. Ionic Bonds
   2. Covalent Bonds (Polar and Nonpolar)
   3. Electron Dot Notation

B. Inter-Molecular Bonding
   1. Van der Waals Forces
   2. Dipole-dipole
   3. Hydrogen Bonding

Lesson G: Solutions

A. Solutions are Mixtures

1. Homogeneous Mixtures of Solute and Solvent
2. Classification of Solutions

(a) Gaseous Solutions
(b) Liquid Solutions
(c) Solid Solutions

B. The Dissolving Process of Solutions

1. Ions in Solution
2. Molecules in Solution
3. Energy Changes
4. Electrolytes and Nonelectrolytes

C. Concentration of Solutions

1. A Molar Solution
2. A $mol/m^3$ Solution
3. Quantitative Aspects

(a) Preparation of Solutions
(b) Determining Molarity of Solutions
(c) Using Molarity of Solutions in Problem Solving
(d) Dilution of Solutions

Lesson H: Solutions (con't)

D. Solubility and Equilibrium

1. Unsaturated Solutions
2. Saturated Solutions
3. Supersaturated Solutions

E. Factors That Affect Solubility

1. Nature of Solute and Solvent
2. Temperature
3. Pressure

F. Experiment 2 - Bond Types and the Solution Process

Lesson I:

A. Introduction
B. Formulas for Carbon Compounds
C. Saturated Organic Compounds
D. Structural Isomers
E. Unsaturated Organic Compounds
F. Reactivity of Unsaturated Hydrocarbons

Lesson I: (cont'd)

G. Food for Thought
H. Ring Compounds
I. Reactivity of Unsaturated Ring Compounds
J. The Benzene Ring and Aromatic Compounds
K. Functional Groups
L. The Alcohols
M. Organic Acids
N. Esters
O. Additional Information

Lesson J

A. Introduction to Laboratory Work
    1. Preparing for Laboratory Work
    2. Apparatus
    3. Laboratory Safety
B. Laboratory Techniques and Procedures
C. Experiment 3 - Qualitative and Quantitative Observations
D. Graphs
E. Experiment 4 - Melting and Freezing of Solids
F. Tests for Some Common Gases

# Chemistry Upgrading (A-J)

## Booklet of Answers and Solutions

Dear Student:

You will not be required to send in lessons A-J in Chemistry Upgrading. Attached you will find a booklet of answers and solutions. Study these lessons very carefully, work on the exercises and then check your own work by using this booklet.

When you feel you have mastered these lessons you should proceed to doing the Chemistry 30 lessons. However if you have difficulty with some concepts you may call us by the government RITE number in your area, providing you are an Alberta resident.

We wish you success in this course.

Ronald Billey

Ronald Billey
Vice Principal
Math and Science Department

Answers to Chemistry 30 Upgrading

## LESSON A, page 11

1. (a) solid (b) liquid
   (c) gas

2.

| | |
|---|---|
| fixed<br>fixed<br>variable | fixed<br>variable<br>variable |

3. If energy is changed from one form to another, the total amount of energy remains constant.

4. $E = mc^2$

5. This is very open ended. For example:
   (a) solar energy to heat and light energy
   (b) wind energy to electrical energy

## LESSON A, page 25

1. (a) 0.454
   (b) 50 000
   (c) 3 000
   (d) 0.016
   (e) 1.5
   (f) 0.150
   (g) 2 500
   (h) 4 500
   (i) 50 000
   (j) 7 600
   (k) 0.675
   (l) 123
   (m) 0.042

2. $V = l\ w\ h$
   $= (16\ cm)\ (10\ cm)\ (8\ cm)$
   $= \mathbf{1280\ cm^3}$

3. (a) 7 cm × 15 cm
   (b) iii 3 kg
   (c) ii 1 mm

4. $V = \frac{m}{D} = \frac{47.25\ g}{11.34\ g\ /\ cm^3} = \mathbf{4.167\ cm^3}$

5. $m = DV$
$= (7.86\ g/cm^3)\ (14.1\ cm^3) =$ **111 g**

6. $V = (5.5\ cm)\ (4.2\ cm)\ (6.8\ cm)$
$= 157.08\ cm^3$

$$D = \frac{m}{V} = \frac{3.032\ kg}{157.08\ cm^3} = 0.01930\ kg/\ cm^3$$

$$= \mathbf{19\ g/cm^3}$$ It is **gold**.

7. $V = 8.7\ cm^3 - 5.3\ cm^3 = 3.4\ cm^3$

$$D = \frac{m}{V} = \frac{6.75\ g}{3.4\ cm^3} = \mathbf{2.0\ g/cm^3}$$

**LESSON A, page 27**

**EXERCISE 3**

1. (a) 3
(b) 2
(c) 3
(d) 1
(e) 3
(f) 2
(g) 2
(h) 4
(i) $\infty$ infinite
(j) 1

2. (a) 1.83
(b) 0.11
(c) $3.26 \times 10^4$
(d) 89.2
(e) 1.6
(f) 43.43
(g) 15.9
(h) 8.4
(i) 72
(j) 0.4
(k) 190

**EXERCISE 4**

1. (a) $4.48 \times 10^6$
(b) $3.32 \times 10^{-3}$
(c) $2.24 \times 10^1$
(d) $7.5 \times 10^{-6}$
(e) $5 \times 10^3$
(f) $8.2 \times 10^{-5}$

**LESSON A, page 28**

2. Total mass = 5850 g
   In scientific notation = $5.9 \times 10^3$g

3. (a) $2.144 \times 10^6 = 2.1 \times 10^6$
   (b) $1.075 \times 10^4 = 1.1 \times 10^4$
   (c) $9 \times 10^2$
   (d) $8.91 \times 10^1$

**END OF LESSON A**

**LESSON B, page 2**

**EXERCISE 1**

(a) P
(b) C
(c) P
(d) P
(e) C
(f) P
(g) C

(b)
physical

**EXERCISE 2**

(a) P
(b) P
(c) P
(d) C
(e) P
(f) C
(g) C
(h) C
(i) P
(j) C
(k) P
(l) C
(m) C
No

**LESSON B, page 5**

1. 

| Mixtures | Pure Substances |
|---|---|
| watermelon | sugar |
| shovel full of dirt | salt |
| contents of pencil sharpener | distilled water |
| vinegar | mercury |
| milk | |
| flask of air | |

2. 

| homogeneous | heterogeneous |
|---|---|
| A | B |
| E | C |
| G | D |
| H | F |
| | I |

3. Open ended.
   For example (a) wiring
   (b) pots, etc.

4. Open ended
   For example (a) curtains
   (b) furniture, etc.

5. (a) salt
   (b) sugar

**LESSON B, page 7**

chemical
diamond, graphite, etc.

**LESSON B, page 8**

1. 

| elements | compounds |
|---|---|
| magnesium | water |
| oxygen | gasoline |
| zinc | paper |
| nickel | acetic acid |
| hydrogen | glass |
| | carbon monoxide |

2. An element is a substance which cannot be broken down into a simpler substance by chemical means.

3. A compound is a substance consisting of two or more elements chemically combined.

4. Element
It converts to lead by a nuclear reaction.

5. Evaporate the water leaving the salt. Or use some other physical separation.

6. True

7. False

8. No, since the exact proportions of each element are required in a pure substance.

9. Yes, since milk is a mixture.

**LESSON B, page 10**

10. Cu or Al (copper or aluminum)
Al
tungesten (W)
zinc (Zn)
helium (He)
oxygen (O)
uranium (U)
hydrogen (H)

11. Elements always combine in a definite ratio by mass.

12. **21.23** g of oxygen
**0.269** g of oxygen
**11.88** g of oxygen
**0.135** g of oxygen

$$\frac{.269}{.135} = \frac{2}{1}$$ (oxygen ratio)

**yes** (illustrates multiple proportions)

**LESSON B, page 11**

1. hotter
2. No
Yes
3. Yes

**LESSON B, page 12**

1. O, Ca, Mg, Si, Na, H, Al, K, Fe

2. Fe

3. Sb, Br, Cu, Mn, Pt, U, As, Cr, Au, Hg, Ag, Zn, Ba, Co, I, Pb, Ni, Sn

**LESSON B, page 13**

1. (a) Determine density or weight
   (b) check conductivity

2. Place them in water. The less dense rocks will float.

3. Dissolve the salt
   Filter out the sand
   Dry the sand
   Evaporate water from filtrate to leave salt

**END OF LESSON B**

**LESSON C, page 1**

The black box experiment is very open ended. However, be sure that the interpretations correlate with the observations. For example, if you "observe" that the object rolls rather than slides, your interpretation of that object would be that it is round. Similarly, be sure all other interpretations correlate with your "observations."

**LESSON C, page 3**

1. Be sure that your conclusions about the objects can be deduced from your observations and interpretations.

**LESSON C, page 3**

2. With a black box experiment, you are working with unknowns. You make hypothesis, observations and interpretations based on indirect evidence.

3. Yes, since you have little direct evidence for your hypothesis and theories.

**LESSON C, page 4**

Other tests could possibly be x-rays or ultrasound.

**LESSON C, page 16**

1. (a) isotopes
   (b) atomic mass or mass number
   (c) protons
   (d) neutrons
   (e) atomic mass
   (f) atomic number
   (g) cathode rays
   (h) nucleus
   (i) electrons
   (j) repel attract
   (k) (i) alpha (ii) beta (iii) gamma
   (l) more
   (m) protons
   (n) 1840
   (o) 1

2. (a) cathode to anode
   (b) cathode rays
   (c) They moved a paddle wheel.
   (d) either a magnet or electric charges
   (e) They are negatively charged particles.

3. Most alpha particles went through the gold foil, but a few were reflected back at very large angles.

**LESSON C, page 18**

4. Alpha particles are positively charged with a mass four times that of hydrogen.

   Beta particles are negatively charged and they are electrons. (Incidently, beta particles originate from the nucleus and not from the electron shells.)

   Gamma rays are electromagnetic waves like light but with a shorter wavelength.

5. The positive charge due to the protons in the nucleus is balanced by the negatively charged electrons around the nucleus.

6. The atom does not collapse due to the centrifugal force as a result of the rotation of the electrons about the nucleus.
Electrons are confined to specific energy levels. The reference on page 10 regarding centrifugal force is a bit simplistic but it would be beyond the scope of this course to go into more detail here.

7. neutrino, meson, positron

8. Isotopes of an element contain the same number of protons but different numbers of neutrons. The atomic masses of different isotopes vary since the number of neutrons vary.

9. (b)

| | | | | |
|---|---|---|---|---|
| | $_{17}Cl^{35}$ | 17p<br>$17e^-$<br>18n | $_{17}Cl^{37}$ | 17p<br>$17e^-$<br>20n |
| (c) | $_{1}H^{1}$ | 1p<br>$1e^-$<br>On | $_{1}H^{3}$ | 1p<br>$1e^-$<br>2n |
| (d) | $_{92}U^{235}$ | 92p<br>$92e^-$<br>143n | $_{92}U^{238}$ | 92p<br>$92e^-$<br>146n |

10. (a) This emphasizes the recurrence of similar properties.

(b) Properties recur periodically throughout the list of elements when they are arranged in order of increasing atomic number.

(c) They are all gases at O°C and normal air pressure. They have virtually no chemical reactivity.

(d) In the upper right hand corner.

(e) decrease

(f) They conduct heat and electricity.
They have a shiny luster.

**LESSON C, page 22**

1. John Dalton
2. monatomic
3. simple whole numbers
4. (i) physical
5. weight relationships
6. atoms
7. the same
8. the same
9. different
10. one to one
11. diatomic

**LESSON C, page 27**

1. (a) See the bottom of page 23

   (b) (i) $2H_2O_{(g)} \rightarrow 2H_{2(g)} + O_{2(g)}$

   (ii) $2H_{2(g)} + O_{2(g)} \rightarrow 2H_2O_{(g)}$

2. See the middle of page 24.

3. very large

4. See 1, 2, and 4 of the summary at the bottom of page 25.

**LESSON C, page 28**

5. (a) $6.02 \times 10^{23}$

   (b) $n = \left(\frac{1g}{32g/mol}\right)\left(6.02 \times 10^{23} \text{ molecules/mol}\right)$

   $= 1.88 \times 10^{22}$ molecules

   or **$2 \times 10^{22}$ molecules** to 1 significant digit

6. $N_{2\,(g)} + 3H_{2\,(g)} \rightarrow 2NH_{3\,(g)}$

   $20\ cm^3 + (80-20)\ cm^3 \rightarrow 40\ cm^3$
   $20\ cm^3 + 60\ cm^3 \rightarrow 40\ cm^3$
   $1 : 3 : 2$

7. $2\ CO_{(g)} + O_{2(g)} \rightarrow 2CO_{2(g)}$
   $100\ cm^3 + 50\ cm^3 \rightarrow 100\ cm^3$
   $2 : 1 : 2$

**END OF LESSON C**

**LESSON D, page 7**

1. (a)

| | | |
|---|---|---|
| $Ag^+$ | $S^{2-}$ | $Ag_2S$ |
| $K^+$ | $I^-$ | KI |
| $Cu^+$ | $S^{2-}$ | $Cu_2S$ |
| $Al^{3+}$ | $S^{2-}$ | $Al_2S_3$ |
| $Ca^{2+}$ | $Cl^-$ | $CaCl_2$ |
| $Fe^{3+}$ | $O^{2-}$ | $Fe_2O_3$ |
| $Zn^{2+}$ | $Br^-$ | $ZnBr_2$ |
| $Ba^{2+}$ | $Cl^-$ | $BaCl_2$ |
| $Li^+$ | $F^-$ | LiF |
| $Na^+$ | $O^{2-}$ | $Na_2O$ |
| $Ca^{2+}$ | $I^-$ | $CaI_2$ |
| $Fe^{2+}$ | $O^{2-}$ | FeO |

(b)

| | | |
|---|---|---|
| $Ba^{2+}$ | $NO_3^-$ | $Ba(NO_3)_2$ |
| $Al^{3+}$ | $OH^-$ | $Al(OH)_3$ |
| $Hg^+$ | $NO_3^-$ | $HgNO_3$ |
| $Fe^{2+}$ | $SO_4^{2-}$ | $FeSO_4$ |
| $Fe^{3+}$ | $SO_4^{2-}$ | $Fe_2(SO_4)_3$ |
| $Li^+$ | $OH^-$ | LiOH |
| $NH_4^+$ | $CO_3^{2-}$ | $(NH_4)_2CO_3$ |
| $Ca^{2+}$ | $C_2H_3O_2^-$ | $Ca(C_2H_3O_2)_2$ |
| $Al^{3+}$ | $HSO_3^-$ | $Al(HSO_3)_3$ |
| $Na^+$ | $CN^-$ | NaCN |
| $Cr^{2+}$ | $PO_4^{3-}$ | $Cr_3(PO_4)_2$ |

**LESSON D, page 13**

1. (a) $Na_2O_2$
   (b) $H_2S_{(aq)}$
   (c) $H_2SO_{4\ (aq)}$
   (d) $H_2SO_{3\ (aq)}$
   (e) $HClO_{4\ (aq)}$
   (f) $CO_{(g)}$
   (g) $P_2O_3$

2. (a) hydriodic acid or hydrogen iodide
   (b) sodium iodide
   (c) potassium iodate
   (d) iodic acid
   (e) hydrogen peroxide
   (f) mercury II oxide
   (g) magnesium sulfate
   (h) magnesium sulfide
   (i) aluminum chloride
   (j) potassium cyanide
   (k) potassium perchlorate
   (l) sodium bisulfite
   (m) iron III hydroxide
   (n) cobalt II chloride
   (o) zinc carbonate

**LESSON D, page 14**

1. (b) h
   (c) e
   (d) a
   (e) c
   (f) k
   (g) g
   (h) i

2. (a)

| | |
|---|---|
| $Hg_2O$ | mercury I oxide |
| HgO | mercury II oxide |
| FeO | iron II oxide |
| $Fe_2O_3$ | iron III oxide |
| SnO | tin II oxide |
| $SnO_2$ | or $Sn_2O_4$ tin IV oxide |

   (b)

| | |
|---|---|
| CO | carbon monoxide |
| $CO_2$ | carbon dioxide |
| $P_2O_3$ | diphosphorus trioxide |
| $P_2O_5$ | diphosphorus pentoxide |

3.

| | |
|---|---|
| $CaSeO_4$ | calcium selenate |
| $FeSeO_4$ | iron II selenate |
| $Fe_2(SeO_4)_3$ | iron III selenate |
| $SnSeO_4$ | tin II selenate |

4. (a) hypophosphorous acid
   (b) lead II hydroxide
   (c) carbon monoxide
   (d) carbon dioxide
   (e) ammonia
   (f) ammonium bromide
   (g) phosphorus pentabromide
   (h) phosphorus trioxide
   (i) ammonium sulfite
   (j) disodium phosphate

**LESSON D, page 19**

1. $O_2$, Ar, Kr, $N_2$

2. $2Ca + O_2 \rightarrow 2CaO$

3. carbon 12

4. An atomic mass unit is the mass of $\frac{1}{12}$ of the carbon 12 atom. A sulfur atom has a mass $\frac{32}{12}$ times the mass of the carbon 12 atom.

5. (a) $6.02 \times 10^{23}$ atoms

(b) 1 mol of N atoms is 14 g

$$\therefore \frac{56\text{g}}{14\text{g/mol}} = \textbf{4 mol} \text{ of atoms}$$

6. 1 mol He = 4.00 g

$\therefore$ m = nM
= (0.30 mol) (4.00 g/mol)
= **1.2 g**

7. (b) M = 65.38 g/mol
m = nM = (1.6 mol) (65.38 g/mol)
= 104.61 g
= **100 g**

(c) M = 35.45 g/mol
m = (4.5 mol) (35.45 g/mol)
= 159.53 g
= **160 g**

8. (a)
40.08
+ 2(14.01) = 28.02
+ 6(16.00) = 96.00
**104.10 g/mol**

(b)
2 × 54.94 = 109.88
5 × 16.00 = 80.00
**189.88 g/mol**

(c)
1 × 55.85 = 55.85
6 × 12.01 = 72.06
9 × 1.01 = 9.09
6 × 16.00 = 96.00
**233.00 g/mol**

9. $X = 6.02 \times 10^{23}$

10. (a) $2(6.02 \times 10^{23}) = 1.204 \times 10^{24}$

(b) $0.5(6.02 \times 10^{23}) = 3.01 \times 10^{23}$

(c) Yes $6(6.02 \times 10^{23}) = 3.612 \times 10^{24}$

(d) $(.01)(6.02 \times 10^{23}) = 6.02 \times 10^{21}$

**END OF LESSON D**

**LESSON E, page 2**

Below are the unbalanced equations for this exercise

1. $Cu + O_2 \rightarrow CuO$
2. $C_8H_{18} + O_2 \rightarrow CO_2 + H_2O$
3. $Fe + O_2 \rightarrow FeO$
4. $Na + H_2O \rightarrow H_2 + NaOH$
5. $K + Br_2 \rightarrow KBr$
6. $Mg + O_2 \rightarrow MgO$
7. $N_2 + H_2 \rightarrow NH_3$
8. $Cl_2 + Na \rightarrow NaCl$

**LESSON E, page 3**

**EXERCISE 2**

$KClO_3$ + energy $\rightarrow$ KCl + $O_2$ (unbalanced)

**LESSON E, page 4**

**EXERCISE 3**

$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

**LESSON E, page 5**

**Yes** (atoms are conserved)
**Yes** (moles of atoms are balanced)

**LESSON E, page 6**

**EXERCISE 4**

$O_2 + Cl_2 \rightarrow 2\ Cl_2O$

**LESSON E, page 7**

$O_2 + 2Cl_2 \rightarrow 2Cl_2O$

atoms conserved? **Yes**

| Left side | element | Right side |
|---|---|---|
| 2 | O | 2 |
| 4 | Cl | 4 |
| total 6 | | total 6 |

**EXERCISE 5**

$C_2H_2 + O_2 \rightarrow 2CO_2 + H_2O$

$C_2H_2 + \frac{5}{2} O_2 \rightarrow 2CO_2 + H_2O$

Multiply the equation by **2**

$\mathbf{2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O}$

atoms conserved? **yes**

**LESSON E, page 8**

(a) $2C_2H_2 + 5O_2 \rightarrow 4CO_2 + 2H_2O$
(b) $2HgO \rightarrow 2Hg + 1O_2$
(c) $8Cu + 1S_8 \rightarrow 8CuS$
(d) $1Cl_2 + 2NaI \rightarrow 1I_2 + 2NaCl$
(e) $2NaOH + 1H_2SO_4 \rightarrow 2H_2O + 1Na_2SO_4$
(f) $3Li + 1AlCl_3 \rightarrow 1Al + 3LiCl$
(g) $1FeCl_3 + 3KOH \rightarrow 1Fe(OH)_3 + 3KCl$
(h) $1C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2$
(i) $2NaCl \rightarrow 2Na + 1Cl_2$
(j) $2K + 2H_2O \rightarrow 1H_2 + 2KOH$

**Note:** in general, the coefficients "1" are omitted.

**LESSON E, page 9**

$2NaOH + 1H_2SO_4 \rightarrow 2H_2O + 1Na_2SO_4$

| | | | |
|---|---|---|---|
| 2 mol NaOH | = **80.00** g | 2 mol $H_2O$ | = **36.04** g |
| 1 mol $H_2SO_4$ | = **98.08** g | 1 mol $Na_2SO_4$ | = **142.04** g |
| | **178.08** g | | **178.08** g |

mass is conserved? **Yes**

**LESSON E, page 10**

1. $C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2$
   $44.11\ g + 160.00\ g \rightarrow 72.08\ g + 132.03\ g$
   $204.11\ g \rightarrow 204.11\ g$

2. $2NaCl \rightarrow 2Na + Cl_2$
   $116.88\ g \rightarrow 45.98\ g + 70.90\ g$
   $116.88\ g \rightarrow 116.88\ g$

3. $2K + 2H_2O \rightarrow H_2 + 2KOH$
   $78.20\ g + 36.04\ g \rightarrow 2.02\ g + 112.22\ g$
   $114.24\ g \rightarrow 114.24\ g$

**LESSON E, page 11**

1. Yes, atoms are never created or destroyed in chemical reactions.
2. No, there are two moles reacting to produce three moles of products.

**LESSON E, page 13**

1. (a) $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

   (b) 2 mol $H_2O$

   (c) 0.125 mol

   (d) $\dfrac{4.00\ g}{16.05\ g/mol} = 0.25$ mol $CH_4$

   $\therefore$ **0.50 mol** $H_2O$ is produced (2:1 ratio)

2. (a) $1CH_4 + 2O_2 \rightarrow 1CO_2 + 2H_2O$

   (b) **1** mol + **2** mol → **1** mol + **2** mol

   (c) X = **0.125** mol of $CO_2$
   produces **0.125** mol of $CO_2$

   $\dfrac{4.00\ g}{\mathbf{16.05}\ g/mol} = 0.25$ mol

   **2** mol of water vapor
   **0.50** mol water vapor

**LESSON E, page 14**

**EXERCISE 10**

1. (a) $\frac{150\ g}{100.09\ g/mol} = \mathbf{1.500\ mol}$

   (b) 1.500 mol (1:1 ratio)

   (c) (1.500 mol) (44.01 g/mol) = **66.02 g**

2. (a) $\frac{4.0\ g}{2.02\ g/mol} = \mathbf{2\ mol}$

   (b) 2 mol $H_2$ requires 2 mol Ca

   m = (2 mol) (40.08 g/mol) = **80.16 g**

3. (a) $2\ Mg + O_2 \rightarrow 2MgO$

   (b) $\frac{48.6\ g}{24.31\ g/mol} = 2.00$ mol Mg

   2 mol Mg → **2 mol MgO**

4. $2NaI + Cl_2 \rightarrow 2\ NaCl + I_2$

   $\frac{63.5\ g}{253.8\ g/mol} = 0.250$ mol $I_2$

   $\therefore$ 0.250 mol $Cl_2$ is required

   m = (0.250 mol) (70.90 g/mol) = **17.73 g**

5. $Pb(NO_3)_2 + 2NaI \rightarrow 2NaNO_3 + PbI_{2\ (s)}$

   Convert 33.1 g of $Pb(NO_3)_2$ into moles:

   $\frac{33.1\ g}{331.21\ g/mol} = 0.100$ mol

   From your **balanced** equation you will note that one mole $Pb(NO_3)_2$ yields one mole $PbI_2$.

   Therefore 0.100 mol $Pb(NO_3)_2$ yields 0.100 mol $PbI_2$

   Molar mass of $PbI_2$ = 460.99 g/mol

   $\therefore$ m = (0.100 mol)(460.99 g/mol)

   = 46.099 g

   (46.1 g correct to 3 significant digits)

**Alternate method** using mass ratios.)

$$\frac{33.1\text{ g}}{331.21\text{ g}} : \frac{x\text{ g}}{460.99\text{ g}}$$

**x = 46.1 g**

6. (a) (i) First write a balanced equation for this reaction:

$$Cl_2 + H_2 \rightarrow 2HCl$$

(ii) Convert 71.0 g of chlorine into moles:

$$\frac{71.0\text{ g}}{71.0\text{ g/mol}} = 1\text{ mol}$$

Convert 10.0 g of hydrogen into moles:

$$\frac{10.0\text{ g}}{2.0\text{ g/mol}} = 5\text{ mol}$$

(iii) From your **balanced** equation you will note that **1** mol $Cl_2$ needs **1** mol $H_2$ to give **2** mol HCl. Hence chlorine limits the reaction; or 1 mol of $Cl_2$ yields 2 mol of HCl. Hydrogen is in **excess**.

(iv) Convert 2 mol of HCl into grams:

$$2\text{ mol} \times 36.5\text{ g/mol} = 73.0\text{ g of HCl}$$

(b) As explained above, you will note that there are: 5 mol − 1 mol = 4 mol of hydrogen gas in excess or 8 grams of hydrogen gas.

**LESSON E, page 17**

***EXERCISE* 11**

$$N_2 + 3H_2 \rightarrow 2NH_3$$

7 kg of nitrogen equals 7000 g

1 mol of $N_2$ equals 28 g

Therefore 7000 g of $N_2$ equals: $\frac{7000\text{ g}}{28\text{ g/mol}} = 250\text{ mol}$

From your balanced equation above you will note that 1 mol $N_2$ reacts with 3 mol $H_2$.

Hence, 250 mol of $N_2$ should react with $250 \times 3 = 750$ mol of $H_2$.

But 1 mol of $H_2$ = 2 g.

The 750 mol of $H_2$ = 750 mol × 2 g/mol

= 1500 g or 1.5 kg

**Alternate method** (using mass ratios)

$$\frac{7 \text{ kg}}{28 \text{ g}} = \frac{x \text{ kg}}{6 \text{ g}} \text{ (from equation)}$$

$$x = \frac{7 \text{ kg} \times 6 \text{ g}}{28 \text{ g}} = 1.5 \text{ kg of hydrogen}$$

**LESSON E, page 19**

(a) synthesis
(b) replacement or rearrangement
(c) ion - ion
(d) ion - ion Note: the better answer is metal − ion but this was not covered.
(e) decomposition
(f) synthesis
(g) synthesis
(h) replacement

**LESSON F, page 4**

1. (a) A chemical bond is any force which tends to hold atoms together.

   (b) The ionization energy is the energy required to remove the most loosely bound electron from a neutral atom.

   (c) Electronegativity determines the attractive forces between atoms and electrons during chemical bond formation. It can also be thought of as a combination between electron affinity and ionization potential.

   (d) An ionic bond is a chemical bond formed when electrons transfer from one atom to another.

   (e) Unlike atoms share their electrons, but the sharing is unequal. The electronegativity difference is greater than zero and less than 1.8.

   (f) This is a bond formed between two atoms having the same electronegativity.

2. In a stable chemical bond, the compound has less energy than the separate atoms. Also, the atoms acquire the electron configuration of the nearest inert gas.

3. electronegativity

4.

| nonpolar covalent | polar covalent | ionic |
|---|---|---|
| $H_2$ | $H_2O$ | KCl |
| $I_2$ | BrCl | $CaCl_2$ |
| | ZnS | $Al_2O_3$ |
| | AgCl | |

5. (a) upper right
   (b) lower left

6. BeS **S**, BrCl **Cl**, MgO **O**, $Cl_2O$ **O**, $OF_2$ **F**

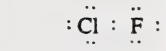
7. F

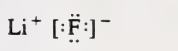
8. (a) decrease
   (b) increase
   (c) increase
   (d) increase
   (e) decrease
   (f) decrease
   (g) larger
   (h) larger
   (i) negatively

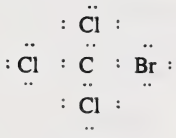
9. The energy of the molecule is less than that of the individual atoms.

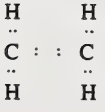
**LESSON F, page 8**

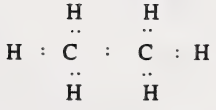
**EXERCISE 2**



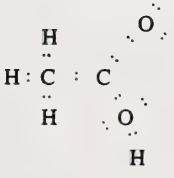








Optional



Below is a summary of the types of bonds that can exist **between** two different molecules. They are listed from the weakest to the strongest.

| | | |
|---|---|---|
| dipole – dipole | (the greater the electronegativity difference the greater the dipole forces) | weakest |
| hydrogen bonding | | |
| ionic bonding | (electronegativity difference is greater than 1.7) | |
| network covalent bonding | e.g. diamond | strongest |

Van der Waals forces due to the number of electrons alone are fairly strong. If no other intermolecular forces are present, the substance with the greatest number of electrons per molecule will have the highest melting and boiling points. Additional forces will increase the melting and boiling points.

Everything else being equal, the stronger the bond type that exists between two different molecules, the higher the melting and boiling points.

1. $PbF_2$ has intermolecular and intramolecular ionic bonds and $PbI_2$ has polar covalent intramolecular bonds which produce dipole – dipole intermolecular forces. So $PbF_2$ must have the highest melting and boiling point since ionic bonding produces stronger intermolecular forces than the dipole – dipole bonding.

2. $Cl_2$ has more electrons, so it has stronger van der Waals forces.

3. Both $O_2$ and NO have about the same van der Waals forces, but different NO molecules also have dipole forces between them, so NO would have the higher melting point.

4. HF has hydrogen bonding between two molecules. HBr has dipole – dipole forces between molecules. Therefore HF has the highest boiling point.

5. (a) (i)
   (b) (ii)

6. T

7. T

8. (a) M
   (b) I
   (c) M
   (d) M
   (e) M

9. $CaCl_2$

**END OF LESSON F**

**LESSON G, page 2**

1. The solute particles:

   (a) are uniformly distributed throughout the solvent
   (b) are molecular in size and cannot be seen
   (c) cannot be removed by filtering
   (d) do not settle out on standing
   (e) form a transparent mixture, whether colorless or colored.

2. (a) solute
   (b) solvent

**LESSON G, page 4**

| | | | | | |
|---|---|---|---|---|---|
| | gas | carbon dioxide | liquid | water | gas in liquid |
| | solid | lemon | liquid | water | solid in liquid |
| alternate for lemonade | liquid | juice | liquid | water | liquid in liquid |

**LESSON G, page7**

3. $KI \rightarrow K^+ + I^-$

4. $CuSO_4 \rightarrow Cu^{2+} + SO_4^{2-}$

5. $MgCl_2 \rightarrow Mg^{2+} + 2Cl^-$

6. $K_3Fe(CN)_6 \rightarrow 3K^+ + Fe(CN)_6^{3-}$

7. $Fe_2(SO_4)_3 \rightarrow 2Fe^{3+} + 3SO_4^{2-}$

8. $NH_4Cl \rightarrow NH_4^+ + Cl^-$

9. $ZnCO_3 \rightarrow Zn^{2+} + CO_3^{2-}$

10. $Ba(NO_3)_2 \rightarrow Ba^{2+} + 2NO_3^-$

**LESSON G, page 10**

1. False
2. An unequal distribution of electrons causes positive and negative ends, or dipoles.
3. The oxygen end, since a positive ion is attracted to the negative end of the water molecule.
4. True
5. True

**LESSON G, page 12**

1. It is ionic since there is a large dissociation into ions.
2. with a large difference

**LESSON G, page 18**

1. $n = CV$

   $n = (0.375 \text{ mol/L})\ (0.200\text{L}) = 0.0750 \text{ mol}$

   $m = nM = (0.0750 \text{ mol})\ (143.32 \text{ g/mol}) = 10.749 \text{ g}$

   $= \mathbf{10.7\ g}$

2. $n = CV$

   $= (0.3 \text{ mol/L})\ (1.000 \text{ L}) = 0.3 \text{ mol}$

   $m = nM = (0.3 \text{ mol})\ (149.89 \text{ g/mol})$

   $= 44.97 \text{ g}$

   $= \mathbf{40\ g}$ (to 1 s.d.)

3. $n = CV$

   $= (0.05 \text{ mol/L})\ (1.5 \text{ L})$

   $= 0.075 \text{ mol}$

   $m = nM$

   $= (0.075 \text{ mol})\ (342.14 \text{ g/mol})$

   $m = 25.66 \text{ g}$

   $\mathbf{m = 30\ g}$

4. $n = \frac{m}{M}$

$= \frac{34.00\ g}{169.88\ g/mol}$

$= 0.200\ mol$

$C = \frac{n}{V}$

$= \frac{0.200\ mol}{0.400\ L}$

$\mathbf{C = 0.500\ mol/L}$

5. $n = \frac{m}{M}$

$= \frac{31.9\ g}{159.61\ g/mol}$

$= 0.19986\ mol$

$C = \frac{n}{V}$

$= \frac{0.19986\ mol}{0.800\ L}$

$\mathbf{C = 0.250\ mol/L}$

**LESSON G, page 19**

6. The balanced equation for this reaction would be:

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

(a) Amount of HCl available = 0.6 L × 0.10 mol/L = 0.06 mol

Since 2 mol HCl → 1 mol $ZnCl_2$ in the reaction above, 0.06 mol HCl will produce **0.03 mol $ZnCl_2$**.

(b) Using the same method as (a) above the 0.06 mol HCl will produce **0.03 mol $H_2$**.

(c) Molar mass of $ZnCl_2$ :

65.38 g/mol
+ 70.90 g/mol
136.28 g/mol

From (a) number of moles = 0.03

m = nM
= (0.03 mol) (136.28 g/mol)
= 4.1 g

7. $K_2CrO_4 + Ba(NO_3)_2 \rightarrow BaCrO_{4(s)} + 2KNO_3$

Convert 126.7 g of barium chromate, $BaCrO_{4(s)}$ into moles.

$\frac{126.7\ g}{253\ g/mol} = 0.5\ mol$

From the equation above you will note that one mole of $BaCrO_4$ requires one mole of $K_2CrO_4$. Now that you know the number of moles of $K_2CrO_4$ required and the volume given in the problem (1 L), you can calculate its molarity.

$M = \frac{mol}{L} = \frac{0.5\ mol}{1\ L} = 0.5\ mol/L$

**LESSON G, page 19**

8. $$C_f = \frac{C_i V_i}{V_f} = \frac{(0.800 \text{ mol/L})(0.200 \text{ L})}{(1.000 \text{ L})} = \mathbf{0.160 \text{ mol/L}}$$

9. $$V_f = \frac{C_i V_i}{C_f} = \frac{(6.000 \text{ mol/L})(0.01500 \text{ L})}{0.250 \text{ mol/L}}$$
$$= \mathbf{0.360 \text{ L}}$$

10. $$V_i = \frac{C_f V_f}{C_i} = \frac{(0.250 \text{ mol/L})(0.600 \text{ L})}{2.000 \text{ mol/L}}$$
$$= \mathbf{0.0750 \text{ L}}$$

**END OF LESSON G**

**LESSON H, page 4**

1. A saturated solution is one in which the solute in solution is in equilibrium with the undissolved solute. (The solute could be solid, liquid, or gas.)

2. Yes, by the operational definition of saturated solution, the undissolved solute is in equilibrium with dissolved solute.

   However, you can decant the solution above the undissolved solute and have a saturated solution with no visible undissolved solute.

3. A supersaturated solution is a solution that contains more dissolved solute than it could normally dissolve at the existing temperature.

4. (a) saturated
   (b) unsaturated
   (c) supersaturated

5. b
6. Dissolved gases are coming out of solution.
7. Oil based paints and varsol are nonpolar.
   Water is polar.
8. increases, weakens
9. more
10. less
11. No
    No
    Yes
12. c
13. a

**LESSON H, page 6**

14. It warms up.
    It decreases.
    It decreases.

15. c
    It depends on how much each is increased, since they have opposite effects.

16. NO – because of the dynamic nature of a saturated solution. Some molecules are entering the solution phase, while others are leaving, hence an **interchange** of molecules is taking place.

**LESSON H, page 7**

*EXERCISE* 2

1. a
2. b
3. c
4. b
5. c
6. It warms up.
   Yes
7. polar
8. remain undissolved
9. No, pressure has virtually no effect on liquids and solids.

**LESSON H, page 8**

10. $Na_2S_2O_3 \cdot 5H_2O_{(aq)} \rightarrow H_2O + Na_2S_2O_{3(s)} + \text{heat}$

11. No

12. The molecular weight of $Na_2S_2O_3 \cdot 5H_2O$ is

$$2(23) + 2(32) + 3(16) + 5(2 + 16)$$

$$= 46 + 64 + 48 + 90$$

$$= 248$$

The molecular weight of $5H_2O$ is $5(2 + 16) = 90$

Therefore $\frac{90}{248}$ is the fraction of water. With 15 g, this turns out to be

$15 \times \frac{90}{248} = 5.44$ g.

**LESSON H, page 9**

| | |
|---|---|
| Yes | No |
| No | Yes |
| Yes | No |
| Yes | No |

**LESSON H, page 10**

1. polar, polar
2. (a) water, sugar, salt
   (b) oil
3. nonpolar, nonpolar
4. Yes, since both are nonpolar.
5. No, since KCl is ionic and gasoline is nonpolar.
6. Yes, since both are nonpolar.

**END OF LESSON H**

**LESSON I, page 19**

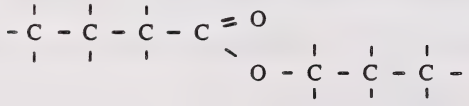
1. carbon

2.

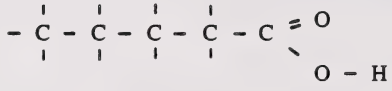
| Fuels | Drugs | Fibers |
|---|---|---|
| wood | aspirin | wool |
| oil | penicillin | nylon |
| gasoline, etc. | tranquilizers, etc. | linen, etc. |

3. (a) hexane
   (b) 2−pentene
   (c) propyne
   (d) propanol
   (e) pentanoic acid
   (f) ethyl butanoate
   (g) 1−ethyl−4−propylbenzene

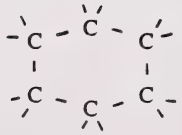
4. (a)



(b)

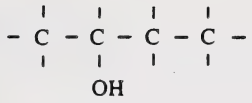


(c)

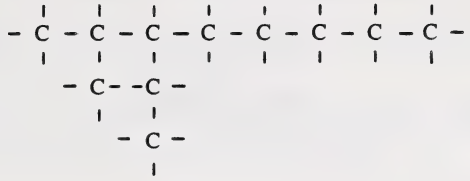


**LESSON I, page 21**

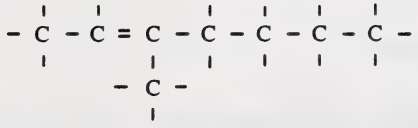
4. (d)



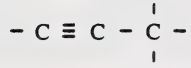
(e)



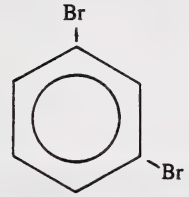
(f)



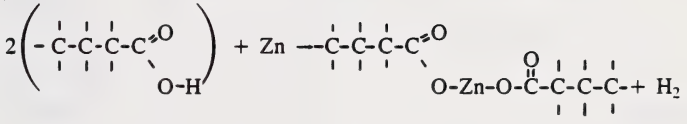
(g)



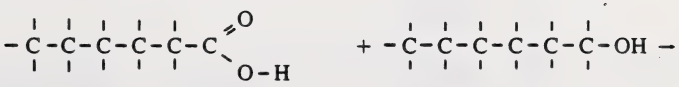
(h)

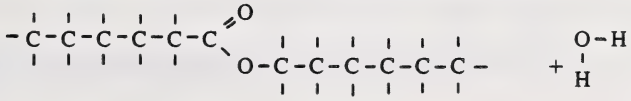



5. (a) — zinc butanoate + hydrogen

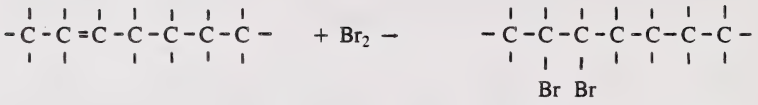


(b) — hexyl hexanoate + water





(c) → 2, 3 dibromoheptane




**END OF LESSON I**

**LESSON J, page 4**

1. plastic or paper
2. meniscus
3. the horizontal line of sight
4. (a) estimate required quantities
   (b) never return substances to stock bottles
   (c) use a clean, dry spatula to transfer solids.
   You may also have others such as placing the cap with the top down.
5. evaporating the liquid
6. This means melting the glass edges to make them smooth.
7. It prevents cuts and makes stopper insertion easier.

**LESSON J, page 6**

In this experiment, the hot water should dissolve the sugar cube the fastest. It should dissolve second fastest in the "crushed-cold", and slowest in the cold water. Also, the whole cubes should show bubbles being released in both the hot and cold water.

**LESSON J, page 7**

1. (a) qualitative
   (b) quantitative
   (c) qualitative
   (d) quantitative
   (e) qualitative
   (f) quantitative

2. (a) Yes
   (b) No
   (c) Trapped air is released as the sugar cube breaks down.

**LERSSON J, page 10**

1. In the **heating curve** you will note a gradual rise in the graph as the compound is heated; this is followed by a plateau where no temperature change occurs. Then the graph rises again.

   In the **cooling curve**, the graph gradually falls as the compound is cooled, followed by a plateau where no temperature change occurs. Then the graph falls again.

2. In heating, the temperature increases up to the melting point; it remains constant until it is all melted; and then it increases again.

   In cooling, the temperature falls until the freezing point; it remains constant until the substance is a solid; then the temperature decreases again.

3. O°C
4. O°C
5. There would be more gradual slopes and longer plateaus.
6. (a) Yes
   (b) Energy is used or released during a phase change and no temperature change occurs.
7. It is pure since there is only one melting and freezing point.

**EXERCISE 5**

1. (a) interpretation
   (b) observation
   (c) observation
   (d) observation
   (e) interpretation

2. (a) qualitative
   (b) quantitative
   (c) qualitative
   (d) interpretation
   (e) 20 $\mu g/m^3$

**LESSON J, page 13**

1. To test for hydrogen, insert a burning splinter into a small sample of gas. If it explodes with a popping sound, it is hydrogen.

2. To test for oxygen, insert a glowing splinter into the gas and if it bursts into flame, it is oxygen.

3. To test for carbon dioxide, place a drop of limewater into the gas. If it turns milky, it is carbon dioxide.

**END OF LESSON J**

# Chemistry Upgrading
# Lesson A

**DO NOT** send this lesson for correction

## TABLE OF COMPLEX IONS

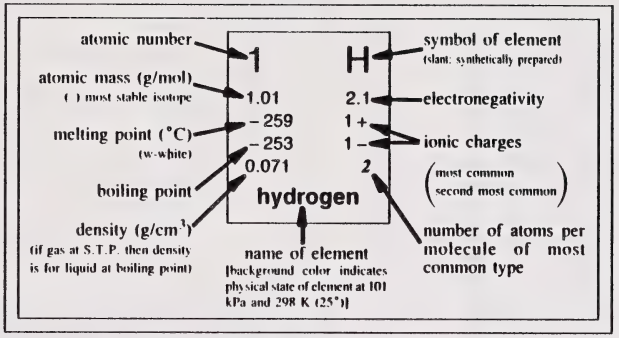
| | | | |
|---|---|---|---|
| acetate | $CH_3COO^-$ | glutamate | $C_5H_8NO_4^-$ |
| ammonium | $NH_4^+$ | hydrogen phosphate | $HPO_4^{2-}$ |
| benzoate | $C_6H_5COO^-$ | hydroxide | $OH^-$ |
| bicarbonate (hydrogen carbonate) | $HCO_3^-$ | hypochlorite | $ClO^-$ |
| bisulfate (hydrogen sulfate) | $HSO_4^-$ | iodate | $IO_3^-$ |
| bisulfide (hydrogen sulfide) | $HS^-$ | nitrate | $NO_3^-$ |
| bisulfite (hydrogen sulfite) | $HSO_3^-$ | nitrite | $NO_2^-$ |
| borate | $BO_3^{3-}$ | oxalate | $OOCCOO^{2-}$ |
| bromate | $BrO_3^-$ | perchlorate | $ClO_4^-$ |
| carbonate | $CO_3^{2-}$ | permanganate | $MnO_4^-$ |
| chlorate | $ClO_3^-$ | phosphate | $PO_4^{3-}$ |
| chlorite | $ClO_2^-$ | silicate | $SiO_3^{2-}$ |
| chromate | $CrO_4^{2-}$ | stearate | $C_{17}H_{35}COO^-$ |
| cyanide | $CN^-$ | sulfate | $SO_4^{2-}$ |
| dichromate | $Cr_2O_7^{2-}$ | sulfite | $SO_3^{2-}$ |
| dihydrogen phosphate | $H_2PO_4^-$ | tetraborate | $B_4O_7^{2-}$ |
| ferricyanide | $Fe(CN)_6^{3-}$ | thiocyanate | $SCN^-$ |
| ferrocyanide | $Fe(CN)_6^{4-}$ | thiosulfate | $S_2O_3^{2-}$ |
| | | tripolyphosphate | $P_3O_{10}^{5-}$ |

## SOLUBILITY OF SOME IONIC COMPOUNDS IN WATER AT 298 K (25°C)

| ION | Group IA: alkali metal ions | Group IA: $H^+$ ($H_3O^+$) | $NH_4^+$ | $NO_3^-$ | $CH_3COO^-$ | $Cl^-$, $Br^-$, $I^-$ | $SO_4^{2-}$ | $S^{2-}$ | $OH^-$ | $PO_4^{3-}$, $SO_3^{2-}$, $CO_3^{2-}$ |
|---|---|---|---|---|---|---|---|---|---|---|
| **SOLUBILITY >0.1 mol/L (VERY SOLUBLE)** | All | All | All | All | Most | Most | Most | Group IA, Group IIA, $NH_4^+$ | Group IA, $NH_4^+$, $Sr^{2+}$, $Ba^{2+}$, $Tl^+$ | Group IA, $NH_4^+$ |
| **SOLUBILITY <0.1 mol/L (SLIGHTLY SOLUBLE)** | None | None | None | None | $Ag^+$ | $Ag^+$, $Pb^{2+}$, $Hg_2^{2+}$, $Cu^+$, $Tl^+$ | $Ca^{2+}$, $Sr^{2+}$, $Ba^{2+}$, $Ra^{2+}$, $Ag^+$, $Pb^{2+}$ | Most | Most | Most |

# Periodic Table of Elements

## KEY




elements do not form ionic compounds

gas

liquid

solid

GROUP

| IA | IIA | IIIB | IVB | VB | VIB | VIIB | VIIIB | VIIIB | VIIIB | IB | IIB | IIIA | IVA | VA | VIA | VIIA | VIIIA |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 1 H<br>1.01 2.1<br>– 259 1+<br>– 253 1–<br>0.071 2<br>hydrogen | | | | | | | | | | | | | | | | | 2 He<br>4.00 –<br>– 272<br>– 269<br>0.126<br>helium |
| 3 Li<br>6.94 1.0<br>180 1+<br>1347<br>0.53<br>lithium | 4 Be<br>9.01 1.5<br>1278 2+<br>2970<br>1.85<br>beryllium | | | | | | | | | | | 5 B<br>10.81 2.0<br>2079<br>2550<br>2.34<br>boron | 6 C<br>12.01 2.5<br>3550<br>4827<br>2.26<br>carbon | 7 N<br>14.01 3.0<br>– 210 3–<br>– 196<br>0.81 *2*<br>1.14 *2*<br>nitrogen | 8 O<br>16.00 3.5<br>– 218 2–<br>– 183<br>1.14 *2*<br>oxygen | 9 F<br>19.00 4.0<br>– 220 1–<br>– 188<br>1.505 *2*<br>fluorine | 10 Ne<br>20.18 –<br>– 249<br>– 246<br>1.20<br>neon |
| 11 Na<br>22.99 0.9<br>97.8 1+<br>883<br>0.97<br>sodium | 12 Mg<br>24.31 1.2<br>649 2+<br>1090<br>1.74<br>magnesium | | | | | | | | | | | 13 Al<br>26.98 1.5<br>660 3+<br>2467<br>2.70<br>aluminum | 14 Si<br>28.09 1.8<br>1410<br>2355<br>2.33<br>silicon | 15 P<br>30.97 2.1<br>44.1 w 3–<br>280 w *3*<br>1.82 w *4*<br>phosphorus | 16 S<br>32.06 2.5<br>113 2–<br>445<br>2.07 *8*<br>sulfur | 17 Cl<br>35.45 3.0<br>– 101 1–<br>– 35<br>1.56 *2*<br>chlorine | 18 Ar<br>39.95 –<br>– 189<br>– 186<br>1.40<br>argon |
| 19 K<br>39.10 0.8<br>63.7 1+<br>774<br>0.86<br>potassium | 20 Ca<br>40.08 1.0<br>839 2+<br>1484<br>1.55<br>calcium | 21 Sc<br>44.96 1.3<br>1541 3+<br>2831<br>3.0<br>scandium | 22 Ti<br>47.90 1.5<br>1660 4+<br>3287 3+<br>4.54<br>titanium | 23 V<br>50.94 1.6<br>1890 5+<br>3380 4+<br>6.1<br>vanadium | 24 Cr<br>52.00 1.6<br>1857 3+<br>2672 2+<br>7.19<br>chromium | 25 Mn<br>54.94 1.5<br>1245 2+<br>1962 4+<br>7.43<br>manganese | 26 Fe<br>55.85 1.8<br>1535 3+<br>2750 2+<br>7.87<br>iron | 27 Co<br>58.93 1.8<br>1495 2+<br>2870 3+<br>8.9<br>cobalt | 28 Ni<br>58.71 1.8<br>1453 2+<br>2732 3+<br>8.9<br>nickel | 29 Cu<br>63.55 1.9<br>1083 2+<br>2567 1+<br>8.96<br>copper | 30 Zn<br>65.38 1.6<br>420 2+<br>907<br>7.13<br>zinc | 31 Ga<br>69.74 1.6<br>29.8 3+<br>2403<br>5.90<br>gallium | 32 Ge<br>72.59 1.8<br>937 4+<br>2830<br>5.32<br>germanium | 33 As<br>74.92 2.0<br>– 7.2 3–<br>615<br>5.73<br>arsenic | 34 Se<br>78.96 2.4<br>217 2–<br>685<br>4.79<br>selenium | 35 Br<br>79.90 2.8<br>– 7.2 1–<br>58<br>3.12 *2*<br>bromine | 36 Kr<br>83.80 –<br>– 157<br>– 152<br>2.41<br>krypton |
| 37 Rb<br>85.47 0.8<br>38.9 1+<br>688<br>1.53<br>rubidium | 38 Sr<br>87.62 1.0<br>769 2+<br>1384<br>2.5<br>strontium | 39 Y<br>88.91 1.3<br>1522<br>3338<br>4.47<br>yttrium | 40 Zr<br>91.22 1.4<br>1852<br>4377<br>6.51<br>zirconium | 41 Nb<br>92.91 1.6<br>2468<br>4742<br>8.6<br>niobium | 42 Mo<br>95.94 1.8<br>2617<br>4612<br>10.2<br>molybdenum | 43 *Tc*<br>98.91 1.9<br>2172<br>4877<br>11.5<br>technetium | 44 Ru<br>101.07 2.2<br>2310<br>3900<br>12.4<br>ruthenium | 45 Rh<br>102.91 2.2<br>1966<br>3727<br>12.4<br>rhodium | 46 Pd<br>106.40 2.2<br>1552 2+<br>3140 4+<br>12.0<br>palladium | 47 Ag<br>107.87 1.9<br>962 1+<br>2212<br>10.5<br>silver | 48 Cd<br>112.41 1.7<br>321 2+<br>765<br>8.65<br>cadmium | 49 In<br>114.82 1.7<br>157 3+<br>2080<br>7.31<br>indium | 50 Sn<br>118.69 1.8<br>232 4+<br>2270 2+<br>7.31<br>tin | 51 Sb<br>121.75 1.9<br>631 3+<br>1750 5+<br>6.70<br>antimony | 52 Te<br>127.60 2.1<br>450 2–<br>990<br>6.24<br>tellurium | 53 I<br>126.90 2.5<br>114 1–<br>184<br>4.93 *2*<br>iodine | 54 Xe<br>131.30 –<br>– 107<br>– 108<br>3.52<br>xenon |
| 55 Cs<br>132.91 0.7<br>28.4 1+<br>678<br>1.90<br>cesium | 56 Ba<br>137.33 0.9<br>725 2+<br>1640<br>3.5<br>barium | 57 La<br>138.91 1.1<br>921<br>3457<br>6.15<br>lanthanum | 72 Hf<br>178.49 1.3<br>2227<br>4602<br>13.3<br>hafnium | 73 Ta<br>180.95 1.5<br>2996<br>5425<br>16.7<br>tantalum | 74 W<br>183.85 1.7<br>3410<br>5660<br>19.3<br>tungsten | 75 Re<br>186.21 1.9<br>3180<br>5627<br>21.0<br>rhenium | 76 Os<br>190.20 2.2<br>3045<br>5027<br>22.6<br>osmium | 77 Ir<br>192.22 2.2<br>2410<br>4130<br>22.5<br>iridium | 78 Pt<br>195.09 2.2<br>1772 4+<br>3827 2+<br>21.5<br>platinum | 79 Au<br>196.97 2.4<br>1064 3+<br>2807 1+<br>19.3<br>gold | 80 Hg<br>200.59 1.9<br>– 39.0 2+<br>357 1+<br>13.6<br>mercury | 81 Tl<br>204.37 1.8<br>304 1+<br>1457 3+<br>11.85<br>thallium | 82 Pb<br>207.19 1.8<br>328 2+<br>1740 4+<br>11.4<br>lead | 83 Bi<br>208.98 1.9<br>271 3+<br>1560 5+<br>9.8<br>bismuth | 84 Po<br>(209) 2.0<br>254 2+<br>962 4+<br>(9.2)<br>polonium | 85 At<br>(210) 2.2<br>302 1–<br>373<br>–<br>astatine | 86 Rn<br>(222) –<br>– 71<br>– 62<br>4.4<br>radon |
| 87 Fr<br>(223) 0.7<br>(27) 1+<br>(677)<br>2.4<br>francium | 88 Ra<br>226.03 0.9<br>700 2+<br>1140<br>5.0<br>radium | 89 Ac<br>(227) 1.1<br>1050 3+<br>3200<br>10.07<br>actinium | 104<br>(261) Rf<br>rutherfordium<br>Ku<br>kurchatovium | 105<br>(262) Ha<br>hahnium<br>Ni<br>nielsbohrium | 106<br>(263) | | | | | | | | | | | | |

| | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| 58 Ce<br>140.12 1.1<br>799 3+<br>3426<br>6.67<br>cerium | 59 Pr<br>140.91 1.1<br>931 3+<br>3512<br>6.77<br>praseodymium | 60 Nd<br>144.24 1.2<br>1021 3+<br>3068<br>7.00<br>neodymium | 61 *Pm*<br>(145) –<br>1080 3+<br>2460<br>7.22<br>promethium | 62 Sm<br>150.35 1.2<br>1077 3+<br>1791<br>7.52<br>samarium | 63 Eu<br>151.96 –<br>822 3+<br>1597 2+<br>5.24<br>europium | 64 Gd<br>157.25 1.1<br>1313 3+<br>3266<br>7.90<br>gadolinium | 65 Tb<br>158.93 1.2<br>1356 3+<br>3123<br>8.23<br>terbium | 66 Dy<br>162.50 –<br>1412 3+<br>2562<br>8.55<br>dysprosium | 67 Ho<br>164.93 1.2<br>1474 3+<br>2695<br>8.80<br>holmium | 68 Er<br>167.26 1.2<br>159 3+<br>2863<br>9.07<br>erbium | 69 Tm<br>168.93 1.2<br>1545 3+<br>1947<br>9.32<br>thulium | 70 Yb<br>173.04 1.1<br>819 3+<br>1194<br>6.97<br>ytterbium | 71 Lu<br>174.97 1.2<br>1663 3+<br>3395<br>9.84<br>lutetium |
| 90 Th<br>232.04 1.3<br>1750 4+<br>4790<br>11.7<br>thorium | 91 Pa<br>231.04 1.5<br>– 1600 5+<br>–<br>15.4<br>protactinium | 92 U<br>238.03 1.7<br>1132 6+<br>3818 4+<br>19.1<br>uranium | 93 *Np*<br>237.05 1.3<br>640 5+<br>3902<br>19.5<br>neptunium | 94 *Pu*<br>(244) 1.3<br>641 4+<br>3232 6+<br>–<br>plutonium | 95 *Am*<br>(243) 1.3<br>994 3+<br>2607<br>13.7<br>americium | 96 *Cm*<br>(247) –<br>1340 3+<br>–<br>–<br>curium | 97 *Bk*<br>(247) –<br>– 3+<br>–<br>14<br>berkelium | 98 *Cf*<br>(251) –<br>– 3+<br>–<br>–<br>californium | 99 *Es*<br>(254) –<br>–<br>–<br>–<br>einsteinium | 100 *Fm*<br>(257) –<br>–<br>–<br>–<br>fermium | 101 *Md*<br>(258) –<br>–<br>–<br>–<br>mendelevium | 102 *No*<br>(259) –<br>–<br>–<br>–<br>nobelium | 103 *Lr*<br>(257) –<br>–<br>–<br>–<br>lawrencium |

CONTENTS

A. Systems of Measurement
B. The Concept of Matter and Energy
C. Density
D. Scientific Notation
E. Significant Figures

## A. Systems of Measurement

The system of measurement that used to be commonly used in Canada and is used in the United States now is the English system where the foot is the fundamental unit of length; the pound is the fundamental unit of mass; the second is the fundamental unit of time and Fahrenheit is the fundamental unit of temperature. This system is not very satisfactory for scientific work and has been replaced by the International System of Units used by scientists throughout the world.

The revised science courses will be conducted using only SI (The International System of Units). There will be no conversions from SI to the Imperial system. As a student, you must realize the importance of proficiency in SI since you will be using the SI units virtually on every page of every lesson. You would not like to have problems in the Chemistry course mainly due to the fact that you are unable to use the units. Therefore apply yourself to master the "Metric System" of measurement.

### SI The International System of Units

**The Change to Metric:** Through the centuries, many measurement systems have developed, evolving from numerous origins, convenient customs, and local adaptations. Most systems have lacked rational structure. The Imperial system - using the yard, quart, and pound - is one such conglomeration of poorly related units.

About 200 years ago, France decided to bring order out of her chaotic measures and the metric system was developed. Although strongly opposed at first, this new system proved effective and gained popularity, so much so, that over 90% of the world's population now lives in countries that have adopted or are changing to the metric system.

**Various Versions of Metric Systems:** There have been several metric systems, but each new version has added more metric units, causing unnecessary complexity. To make matters worse, in some applications there has been a mixture of both Imperial and metric units, and something had to be done to clean house.

**The Latest Version of the Metric System:** In 1960 the International System of Units was established as a result of a long series of international discussions. This modernized metric system, called SI, from the French name, **Le Systeme International d'Unites**, is now, as a general world trend, to replace all former systems of measurement, including former versions of the metric system. Canada has decided to convert to SI.

Many European nations are making the change to SI which is a change from former metric practice. The United Kingdom, Australia, New Zealand, South Africa, and others are adopting SI, while countries such as India, China, and Japan are updating their metric practice to conform to SI. In the United States, major industries are tooling up for metric conversion, and their choice, too, is SI.

It's SI, not S.I.

It's just called SI not the "SI system", since the "S" stands for the word "system".

**SI is Similar but Different:** SI includes familiar metric units such as the metre and kilogram. There are, however, a number of changes from former metric systems. For instance, the centigrade temperature scale is called the Celsius (pronounced sell-cee-us) scale when used for general purposes. This is a change in name only, so that 20°C, formerly read as "twenty degrees centigrade", is now read as "twenty degrees Celsius." There is no change in the scale, only in the name. Water still freezes at 0°C and boils at 100°C (degrees Celsius, that is). This kind of change is not difficult for those who are familiar with older metric systems.

**Numbers moulded metrics:** The metric system was based on the convenience of the decimal number system. Units are related by factors such as 10, 100, and 1000. This makes computation in the metric system much simpler than that with Imperial measures. A great deal of the arithmetic merely involves the shifting of the decimal, without tedious calculations.

**PLEASE:** **Try to think <u>SI Metric</u>** and use these metric measures whenever possible.

**Table of Prefixes**

| Prefix | Symbol | Multiplier | Factor |
|---|---|---|---|
| tera | T | 1 000 000 000 000 | $= 10^{12}$ |
| giga | G | 1 000 000 000 | $= 10^{9}$ |
| mega | M | 1 000 000 | $= 10^{6}$ |
| *kilo | k | 1 000 | $= 10^{3}$ |
| hecto | h | 100 | $= 10^{2}$ |
| deca | da | 10 | $= 10^{1}$ |
| | | 1 | $= 10^{0}$ |
| deci | d | 0.1 | $= 10^{-1}$ |
| *centi | c | 0.01 | $= 10^{-2}$ |
| *milli | m | 0.001 | $= 10^{-3}$ |
| micro | μ | 0.000 001 | $= 10^{-6}$ |
| nano | n | 0.000 000 001 | $= 10^{-9}$ |
| pico | p | 0.000 000 000 001 | $= 10^{-12}$ |
| femto | f | 0.000 000 000 000 001 | $= 10^{-15}$ |
| atto | a | 0.000 000 000 000 000 001 | $= 10^{-18}$ |

* - most commonly used

| Unit | Symbol | Meaning | Example |
|---|---|---|---|
| terametre | Tm | $10^{12}$ m | Distance from sun to Saturn = 1.4 Tm |
| gigametre | Gm | $10^{9}$ m | About 3 times distance from earth to moon. |
| megametre | Mm | $10^{6}$ m | Distance from Calgary to northern Alberta Border. |
| kilometre | km | $10^{3}$ m | Length of brisk 10 min walk. |
| metre | m | 1 m | Height of 3-drawer filing cabinet. |
| millimetre | mm | $10^{-3}$ m | Thickness of a dime. |
| micrometre | µm | $10^{-6}$ m | Size of bacteria. |
| nanometre | nm | $10^{-9}$ m | Length of oil molecule. |
| picometre | pm | $10^{-12}$ m | Wavelength of gamma rays. |
| femtometre | fm | $10^{-15}$ m | Diameter of a proton. |
| attometre | am | $10^{-18}$ m | ??? |

| Quantity | Name | Symbol | Notes |
|---|---|---|---|
| length | millimetre<br>centimetre<br>metre<br>kilometre | mm<br>cm<br>m<br>km | thickness of a dime.<br><br>A Volkswagen is about 4 m long. |
| area | square centimetre<br>square metre | $cm^2$<br>$m^2$ | |
| volume | cubic centimetre<br>cubic metre<br>millilitre<br>litre | $cm^3$<br>$m^3$<br>mL<br>L | 1 cm by 1 cm by 1 cm.<br>measure of gravel, concrete.<br>1000 mL = 1 L; 1 mL = 1 $cm^3$<br>1 L is the volume of a cube 10 cm by 10 cm by 10 cm. |
| mass | gram<br>kilogram<br>milligram<br>tonne | g<br>kg<br>mg<br>t | mass of 1 $cm^3$ of water at 4°C.<br>mass of 1 L of water at 4°C.<br><br>1 t is 1 000 kg, this is about the mass of a Volkswagen. |
| temperature | degree Celsius | °C | A comfortable room has a temperature of 20°C. |
| time | hour<br>minute<br>second | h<br>min<br>s | |
| electric current | ampere | A | |
| electric potential,<br>potential difference,<br>electromotive force | volt | V | |

## RULES FOR WRITING SI

1. The symbols are always printed in Roman (upright) type, irrespective of the type face used in the rest of the text.

2. Symbols are never pluralized: 45 g (not 45 gs)

3. Never use a period after a symbol, except when the symbol occurs at the end of a sentence. This is done because SI symbols are SYMBOLS they are NOT abbreviations. (i.e., kg not kg.)

4. Symbols should usually be used and unit names not mixed with symbols.

   Example: 10 kg (preferred), ten kilograms (accepted), never 10 kilograms.

5. Always use a full space between the quantity and the symbol: 45 g (not 45g)

   Exception: For Celsius temperatures the degree sign occupies the space. 32°C (not 32° C or 32 °C)

6. Use decimals, not fractions: 0.25 g (not $\frac{1}{4}$ g)

7. For values less than one, use a zero in front of the decimal (0.45 g not .45 g).

8. Symbols are written in lower case, except when the unit is derived from a proper name:

   m for metre; s for second; but N for newton; A for ampere; degree Celsius °C is the only one to be upper case in both name and symbol.

9. Prefixes are printed in Roman (upright) type without spacing between the prefix and the unit symbol: kg for kilogram, km for kilometre

   Only one prefix is applied at one time to a given unit: megagram or tonne, NOT kilokilogram

10. Use spaces to separate long lines of digits into easily readable blocks of three digits with respect to the decimal marker: 32 453.246 072 5

    Exception: A space is optional with a four-digit number: 1 234 or 1234

11. Division of Units in symbolic form is indicated by an oblique stroke between the symbols, <u>or</u> by a negative exponent.

12. In general, the term mass replaces the term weight. i.e., do not use the term weight or weighs.

## Metric Equivalents

Page 4 contains a summary of the metric quantities. Page 2 contains a table of prefixes and their respective factors and multipliers. Please refer to these pages as required.

Use your common sense as you work through these sections to ensure your answers are reasonable. An example of this could be illustrated by imagining yourself as a teacher with a pupil who is learning about money. If the student was to give the equivalent of $1.00 in pennies, how would you reply to an answer of 0.01 pennies? How would you reply to the number of dimes in $0.50 as 500? Probably you would reply "Use your common sense! Does your answer seem reasonable?" The same "reasonable answers" apply to your work on metric equivalents or to any work you do in chemistry.

## Length

The base unit of length is the metre. The most commonly used length units are the km, m, cm, and mm.

Equivalents:

| | | | |
|---|---|---|---|
| 1 km = 1000 m | or | 1 m = 0.001 km |
| 1 m = 100 cm | | 1 cm = 0.01 m |
| 1 m = 1000 mm | | 1 mm = 0.001 m |

Examples on finding metric length equivalents:

(a) 5.3 m = _____ cm

since 1 m = 100 cm
5.3 m = (5.3 m)(100 cm/m)
= 530 cm

(b) 35 mm = _____ m

since 1 m = 1000 mm or 1 mm = 0.001 m
35 mm = (35 mm)(0.001 m/mm)
= 0.035 m

(c) 4500 cm = _____ km

since 1 cm = 0.01 m and 1 m = 0.001 km
then 1 cm = 0.000 01 km
4500 cm = (4500 cm)(0.000 01 km/cm)
= 0.045 km

An alternate solution is to do the question in 2 steps:

4500 cm = _____ m

4500 cm = (4500 cm)(0.01 m/cm)
= 45 m

then

45 m = _____ km

45 m = (45 m)(0.001 km/m)
= 0.045 km

Now try all of the practice exercises and check your answers with those at the end of this lesson. (pages 23 and 24) Ensure you can do all questions correctly before proceeding.

## Practice Exercises

1. (a) 143 cm = _____ m
(b) 290 mm = _____ m
(c) 9.10 km = _____ m
(d) 7.95 m = _____ cm
(e) 0.075 m = _____ mm
(f) 7500 m = _____ km
(g) 0.025 km = _____ cm
(h) 1500 mm = _____ cm
(i) 0.03 cm = _____ mm

## Mass

The base unit of mass is the kilogram. The most commonly used mass units are kg, g, and mg.

Note: The term mass replaces the term weight. Mass is the amount of matter in a body.

Equivalents:

1 kg = 1000 g
1 g = 1000 mg

1 mg = 0.001 g
1 g = 0.001 kg

Examples on metric mass equivalents:

(a) 6.7 kg = _____ g

6.7 kg = (6.7 kg)(1000 g/kg)
= 6700 g

(b) 478 mg = _____ g

478 mg = (478 mg)(0.001 g/mg)
= 0.478 g

(c) 9500 mg = _____ kg

9500 mg = (9500 mg)(0.001 g/mg)(0.001 kg/g)
= 0.009 5 kg

**Practice Exercises**

2. (a) 500 g = _____ kg
(b) 0.003 kg = _____ g
(c) 0.007 g = _____ mg
(d) 2.88 mg = _____ g
(e) 0.01 kg = _____ mg
(f) 375 mg = _____ kg

**Volume**

The base unit of volume is the litre. The most commonly used units are L, mL, $cm^3$, and $m^3$.

Equivalents:

1 L = 1000 mL
1 mL = 1 $cm^3$
1 $m^3$ = 1000 L

1 mL = 0.001 L
1 L = 0.001 $m^3$

Note: when multiplying length units to determine area or volume, the length units must all be the same. For example, you cannot multiply m by cm. You can multiply cm by cm or m by m.

Examples on metric volume equivalents:

(a) 43.7 mL = _____ L

43.7 mL = (43.7 mL)(0.001 L/mL)
0.043 7 L

(b) 0.75 L = _____ mL

0.75 L = (0.75 L)(1000 mL/L)
= 750 mL

(c) 0.02 $m^3$ = _____ L

0.02 $m^3$ = (0.02 $m^3$)(1000 L/$m^3$)
= 20 L

(d) 3 m x 5 m x 200 cm = _____ $m^3$

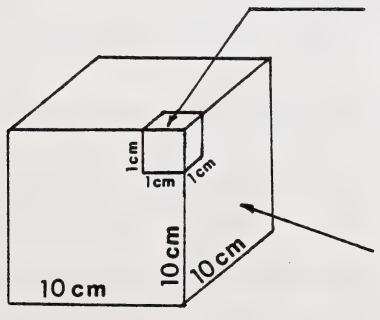
first convert 200 cm into m
200 cm = (200 cm)(0.01 m/cm)
= 2 m
thus 3 m x 5 m x 2 m = 30 $m^3$

**Practice Exercises**

3. (a) 375 mL = _____ L
(b) 0.65 L = _____ $cm^3$
(c) 3.2 $m^3$ = _____ L
(d) 30 cm x 5 mm x 10 mm = _____ $cm^3$
(e) 0.01 m x 2 cm x 35 cm = _____ $cm^3$
(f) 9.8 L = _____ mL

**A Special Relationship** between length, volume, and mass.

The diagram below shows that a cube 1 cm to a side has a volume of 1 $cm^3$ which is equal to 1 mL. This volume of pure water has a mass very close to 1 g (we often refer to the mass of 1 mL of water as 1 g). Thus 1 L of water has a mass of 1 kg.




1 $cm^3$ or 1 mL contains one gram (1 g) of water at 4°C.

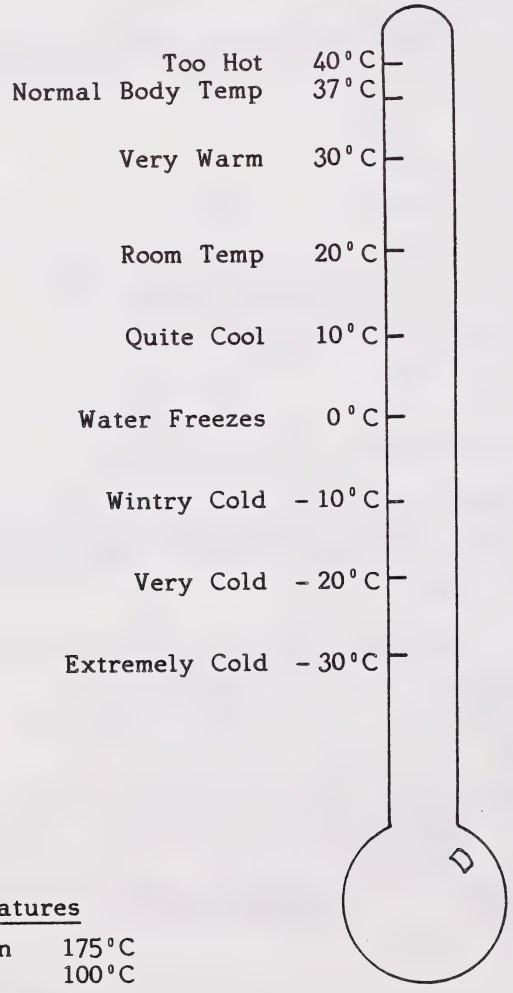
1000 $cm^3$ or one litre contains 1 kg of water at 4°C.

## Time

The basic SI unit for time is the second (s). Due to its universal acceptance, time is not divided into units of ten, so there are still sixty seconds in one minute, sixty minutes in one hour, and twenty-four hours in a day.

## Temperature

In SI, the Celsius scale is used to measure temperature. In this scale there are exactly one hundred units between the freezing temperature of water ($0^{\circ}C$) and its boiling temperature ($100^{\circ}C$). The scale below indicates some common temperatures.




### Other Temperatures

| | |
|---|---|
| Moderate Oven | 175°C |
| Water Boils | 100°C |

## B. The Concept of Matter and Energy

Matter is anything which takes up space and has mass. Matter is generally classified into three forms: solid, liquid, and gas. The volume of a given mass of solid or liquid is constant, but volumes of gases vary greatly with temperature and pressure. The shape of a solid is constant, but the shape of a liquid or gas depends on the shape of the container. In a chemical change it does NOT matter which form the matter is in but the total mass of the reactants is equal to the total mass of the products. This is known as the Law of Conservation of Matter.

Energy, like matter, can also exist in many forms. Some of these forms are heat, light, chemical, electrical, mechanical, sound energy, and others. As in the case with matter, energy can be converted from one form to the next. However, energy is never lost. This is known as the Law of Conservation of Energy.

In unique circumstances, matter can be converted into energy. This occurs in the sun and in nuclear reactions. Albert Einstein first suggested the relationship $E = mc^2$, where E is energy, m is mass, and c is the speed of light. In these circumstances, it appears that mass can be lost to produce a huge amount of energy. Instead of talking about the conservation of matter and the conservation of energy, it would be more correct to talk of the Law of Conservation of Mass-Energy.

**Submit to the Correspondence School for correction and grading all exercises for which answers are not provided.**

### Exercise 1

1. Matter exists in three forms (a) _____, (b) _____, and (c) _____.

2. In the appropriate space in the table below, write in the word "fixed" or "variable".

| | Volume | Shape |
|---|---|---|
| solid | | |
| liquid | | |
| gas | | |

3. What is the Law of Conservation of Energy?

_____

_____

_____

_____

4. Albert Einstein showed the relationship between matter and energy. Write the equation that shows this relationship.

_____

5. From your own experiences, give two examples of how energy is changed from one form to another.

(a) _____

_____

(b) _____

_____

## C. Density

Density is the mass of a substance per unit volume. Mass alone cannot give you the density since the volume is not considered. For example, a kilogram of feathers and a kilogram of gold have the same mass but gold is much denser since it will occupy much less space.

In general, gases are the least dense, liquids are much denser than gases, and solids are the most dense. The states of matter and the relative densities are referred to in general since there are exceptions. A notable exception is water which is more dense in liquid form than in solid form (ice). Were this not true, water would freeze from the bottom up since ice would sink if it was more dense than water. We know that solid water floats on liquid water from personal experience. Most other materials do not behave this way.

The common units of density are $g/cm^3$ and $kg/m^3$. Here is a list of the densities of some common materials:

| Substance | Density |
|---|---|
| water | 1 $g/cm^3$ |
| aluminum | 2.7 $g/cm^3$ |
| iron | 7.86 $g/cm^3$ |
| lead | 11.34 $g/cm^3$ |
| gold | 19.3 $g/cm^3$ |
| mercury | 13.6 $g/cm^3$ |

The formula for density is $D = \frac{m}{V}$

Where D = density
m = mass
V = volume

Mathematical rearrangement yields two formulas:

$m = DV$ and $V = \frac{m}{D}$

**Examples**

(a) An object whose mass is 12.75 g has a volume of 3.20 $cm^3$. Determine the density of this object.

$$D = \frac{m}{V}$$

$$D = \frac{12.75\ g}{3.20\ cm^3} = 3.98\ g/cm^3$$

(b) An object has a density of 5.20 $g/cm^3$ and a volume of 200.0 $cm^3$. Find the mass of the object.

$$m = DV$$

$$m = (5.20\ g/cm^3)(200.0\ cm^3) = 1040\ g.$$

(c) Sulfur has a density of 2.0 $g/cm^3$. Determine the volume of 44 g of sulfur.

$$V = \frac{m}{D}$$

$$V = \frac{44\ g}{2.0\ g/cm^3} = 22\ cm^3$$

**Practice Exercises**

4. Determine the missing quantity.

| | Mass | Volume | Density |
|---|---|---|---|
| (a) | 7.55 g | 30.0 $cm^3$ | _____ |
| (b) | _____ | 5.0 $cm^3$ | 8.0 $g/cm^3$ |
| (c) | 0.50 g | _____ | 0.02 $g/cm^3$ |

## D. Scientific Notation

Scientific notation (also known as standard form) is the method of writing numbers as a number between 1 and 10 multiplied by a power of ten.

Scienfific notation: $y \times 10^x$ where $1 \leq y < 10$ and x is an integer. That is, y must be greater than or equal to one but less than 10. The value y will always have one and only one digit in the ones place.

Examples of scientific notation:

$1 \times 10^2$

$9.9 \times 10^{-3}$

$5.663 \times 10^{-30}$

$6.02 \times 10^{23}$

Examples of numbers NOT in scientific notation:

$35 \times 10^3$

$1.2 \times 10^{-5.5}$

$0.72 \times 10^6$

Scientific notation is used for two main reasons:

(a) writing measurements with the proper number of significant digits.

(b) calculations involving very large and/or very small numbers.

Study the table below.

$1000 = 10^3$ exponent is 3

$100 = 10^2$ exponent is 2

$10 = 10^1$

$1 = 10^0$

$\frac{1}{10} = 0.1 = 10^{-1}$ Note: $10^2 = \frac{1}{10^{-2}}$

$\frac{1}{100} = 0.01 = 10^{-2}$ $10^{-2} = \frac{1}{10^2}$

$\frac{1}{1000} = 0.001 = 10^{-3}$

$0.000\ 000\ 1 = 10^{-7}$

Any number may be expressed as an integral power of ten or as the product of two numbers one of which is an integral power of ten.

e.g. $34\ 500 = 3.45 \times 10^4$

**Rule 1:**

> A shift of the decimal point to the left requires the use of a positive exponent.

e.g. $1985 = 198.5 \times 10^1$

$= 19.85 \times 10^2$

$= 1.985 \times 10^3$

**Rule 2:**

> A shift of the decimal point to the right requires the use of a negative exponent.

e.g. $1.985 = 19.85 \times \frac{1}{10} = 19.85 \times 10^{-1}$

$1.985 = 198.5 \times \frac{1}{10^2} = 198.5 \times 10^{-2}$

$0.000\ 000\ 48 = \frac{4.8}{10\ 000\ 000} = \frac{4.8}{10^7} = 4.8 \times 10^{-7}$

Many measurements in science involve very large or very small numbers which often need to be multiplied or divided. So that this can be done conveniently, such numbers should be expressed as two factors. The first factor known as the digital factor is a number between 1 and 10 or a number with one integer to the left of the decimal point e.g. 2.45, 3.6, 1.0 and 8.007. The other factor is called the exponential factor and is the correct power of 10.

**Examples**

NUMBER = Digital Factor x Exponential Factor

| | | | | |
|---|---|---|---|---|
| 185 | = | 1.85 | x | $10^{2}$ |
| 1850 | = | 1.85 | x | $10^{3}$ |
| 185 000 | = | 1.85 | x | $10^{5}$ |
| 0.185 | = | 1.85 | x | $10^{-1}$ |
| 0.000 185 | = | 1.85 | x | $10^{-4}$ |

Notice that the digital factor in all these numbers is the same. Only the exponential factor varies.

**Practice Exercises**

5. Write in scientific notation.

(a) 300
(b) 0.097 5
(c) 595 000
(d) 0.000 072
(e) $83 \times 10^{3}$
(f) $0.075 \times 10^{-2}$

6. Write as single decimal numbers.

(a) $3.3 \times 10^{4}$
(b) $8.1 \times 10^{-5}$
(c) $5 \times 10^{5}$
(d) $1.1 \times 10^{-3}$
(e) $0.09 \times 10^{5}$
(f) $300 \times 10^{-7}$

**Addition and Subtraction of Exponential Numbers**

> Convert all numbers to the **same power** of ten before carrying out addition or subtraction.

Example: Add $7.62 \times 10^{3}$ cm and $5.2 \times 10^{2}$ cm

$5.2 \times 10^{2}$ cm = $0.52 \times 10^{3}$ cm

Therefore $7.62 \times 10^{3}$ cm

$+ \underline{0.52 \times 10^{3} \text{ cm}}$

$8.14 \times 10^{3}$ cm

**Note:** Regardless of what power of ten you work with, your final answer should be written in scientific notation.

**Practice Exercises**

7. (a) $3.4 \times 10^{6} + 2.5 \times 10^{4}$
   (b) $1.7 \times 10^{-4} + 9.3 \times 10^{-3}$
   (c) $8.7 \times 10^{5} - 9.6 \times 10^{4}$
   (d) $6.2 \times 10^{-21} - 5.9 \times 10^{-22}$

## Multiplication of Exponential Numbers

| Add exponents of like bases |
|---|

Example:

$$(3 \times 10^{2})(5 \times 10^{3}) = (3 \times 5)(10^{2} \times 10^{3})$$
$$= 15 \times 10^{2+3}$$
$$= 15 \times 10^{5}$$
$$= 1.5 \times 10^{6}$$

## Division of Exponential Numbers

| Subtract exponents of like bases |
|---|

Example:

(a) $\frac{3 \times 10^{2}}{5 \times 10^{3}} = \frac{3}{5} \times 10^{2-3}$

$$= 0.6 \times 10^{-1}$$
$$= 6 \times 10^{-2}$$

(b) $\frac{2.5 \times 10^{-5}}{5.0 \times 10^{-7}} = \frac{2.5}{5.0} \times 10^{-5-(-7)}$

$$= 0.5 \times 10^{2}$$
$$= 5 \times 10^{1}$$

**Practice Exercises**

8. (a) $3.2 \times 10^{5} \times 10^{2}$
   (b) $(4.8 \times 10^{-3})(6.6 \times 10^{-4})$
   (c) $\frac{6.02 \times 10^{23}}{3.01 \times 10^{21}}$
   (d) $\frac{9.8 \times 10^{-5}}{4.7 \times 10^{-7}}$

## E. Significant Digits

Suppose we have a measurement of 2867 mm. This measurement could be expressed as 286.7 cm or 28.67 dm or 2.867 m. These measurements contain exactly the same four digits. Only the position of the decimal point differs. The four digits 2, 8, 6 and 7 are called **significant digits**. We consider measurements to be **equally accurate** if they contain the same number of significant digits. The measurements listed above are equally accurate since each has four significant digits.

**Rules for Determining Significant Digits**

1. Any non-zero digit is significant since it is part of the actual measurement.

2. Zero digits may or may not be significant, depending upon their position in the number.

   (a) Any zero digits between non-zero digits **are** significant, because they are part of the actual measurement.

   (b) Zeros on the right in whole numbers **are not** significant (unless information to the contrary is given), since these zeros usually indicate that the measurement has been rounded.

   (c) Zero digits on the right in decimal numbers **are** significant, since they indicate the number is accurate to that number of decimal places.

   (d) Zero digits on the left in decimal numbers **are not** significant, since they only function as placeholders.

**Examples**

Note which zeros are significant and which are not significant in the following numbers.

| | |
|---|---|
| 1003 | These zeros are significant. (See Rule 2 (a)) |
| 3500 | These zeros are **not** significant. (See Rule 2 (b)) |
| 5.487 0 | This zero is significant. (See Rule 2 (c)) |
| 402 | This zero is significant. (See Rule 2 (a)) |
| 0.006 27 | These zeros are **not** significant. (See Rule 2 (d)) |

Practice Exercises

9. Write "yes" if the boldface zeros are significant and "no" if they are not. State which of the rules from page 1 that you are following.

| Number | Significant, Yes or No? | Rule |
|---|---|---|
| 1 . **0** 7 6 | | |
| 3 . 0 7 **0** 4 | | |
| 4 **0** . 8 0 0 | | |
| 0 . **0** 0 5 0 3 | | |
| 6 . **0** 0 | | |
| **0** . 6 9 | | |

There are occasions when it is difficult to determine if the zeros are significant, or not. In the measurement 6600 m, we cannot immediately tell whether these zeros are significant. In such a case, we always consider the zeros to be placeholders, unless some indication of the accuracy of the measurement is given. The accuracy could be indicated by an expression such as 6600 ± 0.5 m which tells us that the measurement is accurate at the 1 m level. In this case both zeros would be significant.

The best way to indicate that zeros are significant is to write the measurement in scientific notation. For example, the number 6600 could be written as $6.600 \times 10^3$. The zeros in the first factor would be considered significant according to Rule 2 (c).

To find how many significant digits there are in any measurement, first decide if any zeros are significant or are not significant. Then, add the number of significant zeros to the number of non-zero digits.

Practice Exercises

10. For each measurement, circle the zeros that **are not** significant. Then state the number of significant digits in each measurement.

| | | No. of Significant Digits |
|---|---|---|
| (a) | 3.560 7 g | _____ |
| (b) | $6.05 \times 10^5$ mL | _____ |
| (c) | 620.8 L | _____ |
| (d) | 48 052.36 $g/cm^3$ | _____ |
| (e) | 0.006 048 m | _____ |
| (f) | 4.306 00 L | _____ |
| (g) | 0.346 800 $cm^2$ | _____ |
| (h) | 0.000 300 kg | _____ |
| (i) | $6.70 \times 10^4$ mm | _____ |
| (j) | $4.03 \times 10^5$ t (tonne) | _____ |
| (k) | $6.000 \times 10^{-2}$ $m^3$ | _____ |
| (l) | 0.047 $cm_e$ | _____ |

## Adding or Subtracting Measured Quantities

When adding or subtracting measured quantities, the calculated answer should be rounded to the same number of **decimal places** as found in the quantity having the least number of decimal places.

Example:

$$\begin{array}{rl} 38.5 & \text{mm (least number of decimal places)} \\ 0.123 & \text{mm} \\ \underline{19.50} & \underline{\text{mm}} \\ 58.123 & \text{mm} \end{array}$$

The answer should be rounded to 58.1 mm.

## Multiplying and Dividing Measurements

When you multiply or divide measurements, your answer should not have more significant digits than any of the original measurements. If you leave your answer with more significant digits than any of the original measurements, it will appear more accurate than it really is. **IMPORTANT RULE: For a given problem you should determine which measurement has the least number of significant digits and then round the final answer to this same number of significant digits. Do not round off individual calculations within the problem, only round the final answer.**

**Example 1:**

Find the area of a sheet of paper which measures 23.60 cm by 6.81 cm.

**Solution**

Since one of the given measurements has 4 significant digits and the other has 3 significant digits, the product should contain only 3 significant digits.

$\ell$ (length) = 23.60 cm (4 significant digits)
w (width) = 6.81 cm (3 significant digits)

$A = \ell w$ (Area = length x width)

$= 23.60 \text{ cm} \times 6.81 \text{ cm}$

$= 160.716 \text{ cm}^2$

$= 161 \text{ cm}^2$

(rounded to the appropriate 3 significant digits)

Thus, the area of the sheet of paper is 161 $\text{cm}^2$.

**Example 2:**

What is the width of fence used for a rectangular yard if the area enclosed is 42.3 $m^2$ and the length is 7.1 m?

**Solution:**

Since $A = \ell w$

$$w = \frac{A}{\ell}$$

$$= \frac{42.3\ m^2}{7.1\ m}$$

(42.3 $m^2$ ← 3 significant digits; 7.1 m ← 2 significant digits)

$$= 5.957\ 7...m$$

$$= 6.0\ m \text{ (to 2 significant digits)}$$

(The zero **must** be included.)

The answer must be rounded to 2 significant digits, since 2 is the least number of significant digits in either factor.

The width of the yard is 6.0 m.

There is a tendency to overuse significant digits. Be reasonable in your answers if you are using a calculator to calculate answers. When using metric conversions or counting numbers, remember that they are exact numbers and have an infinite number of significant digits.

Example: the mass of 3 coins whose average mass is 4.73 g is

3 x 4.73 g = 14.2 g correct to 3 significant digits.

The 3 is an exact number since you do not count, for example, 3.01 coins therefore 3 has an infinite number of significant digits.

**Practice Exercises**

11. Perform the calculations and round off the answers to the correct number of significant figures

(a) Add
$$\begin{array}{r} 3.37\ g \\ 9.475\ g \\ \underline{21.2\ \ g} \end{array}$$

(b) Add
$$\begin{array}{r} 0.05\ L \\ 1.1\ \ L \\ \underline{9.905\ L} \end{array}$$

(c) Subtract
$$\begin{array}{r} 200.0\ m \\ \underline{10.67\ m} \end{array}$$

(d) Subtract
$$\begin{array}{r} 0.055\ kg \\ \underline{0.01\ \ kg} \end{array}$$

(e) Multiply
$(3\ g/mol)(6.91\ mol)$

(f) Multiply
$(0.035\ g)(31^\circ C)(4.19\ J/g^\circ C)$

(g) Divide
$$\frac{12.6\ \text{g}}{12.01\ \text{g/mol}}$$

(h) Divide
$$\frac{2.00\ \text{g}}{0.019\ \text{g/cm}^3}$$

(i) Convert 4.02 g to mg.

## Answers to Practice Exercises in Lesson A

1. (a) 1.43 m
(b) 0.290 m
(c) 9 100 m
(d) 795 cm
(e) 75 mm
(f) 7.5 km
(g) 2 500 cm
(h) 150 cm
(i) 0.3 mm

2. (a) 0.5 kg
(b) 3 g
(c) 7 mg
(d) 0.002 88 g
(e) 10 000 mg
(f) 0.000 375 kg

3. (a) 0.375 L
(b) 650 $cm^3$
(c) 3 200 L
(d) 15 $cm^3$
(e) 70 $cm^3$
(f) 9 800 mL

4. (a) 0.252 $g/cm^3$
(b) 40 g
(c) 25 $cm^3$

5. (a) $3 \times 10^2$
(b) $9.75 \times 10^{-2}$
(c) $5.95 \times 10^5$
(d) $7.2 \times 10^{-5}$
(e) $8.3 \times 10^4$
(f) $7.5 \times 10^{-4}$

6. (a) 33 000
(b) 0.000 081
(c) 500 000
(d) 0.001 1
(e) 9 000
(f) 0.000 03

7. (a) $3.425 \times 10^6$
(b) $9.47 \times 10^{-3}$
(c) $7.74 \times 10^5$
(d) $5.61 \times 10^{-21}$

8. (a) $3.2 \times 10^7$
(b) $3.168 \times 10^{-6}$
(c) $2 \times 10^2$
(d) $2.09 \times 10^2$

9.

| Number | Significant Yes or No? | Rule |
|---|---|---|
| 1 . (0) 7 6 | yes | 2(a) |
| 3 . 0 7 (0) 4 | yes | 2(a) |
| 4 (0) . 8 0 0 | yes | 2(a) |
| 0 . (0 0) 5 0 3 | no | 2(d) |
| 6 . (0 0) | yes | 2(c) |
| (0) . 6 9 | no | 2(d) |

10.

| | MEASUREMENT | sig. digs. | Rule |
|---|---|---|---|
| (a) | 3.560 7 g | 5 | Rule 2(a) |
| (b) | $6.05 \times 10^5$ mL | 3 | Rule 2(a) |
| (c) | 620.8 L | 4 | Rule 2(a) |
| (d) | 48 052.36 g/cm$^3$ | 7 | Rule 2(a) |
| (e) | (0.00)6 048 m | 4 | Rule 2(d) |
| (f) | 4.306 00 L | 6 | Rule 2(a), 2(c) |
| (g) | (0)364 800 cm$^2$ | 6 | Rule 2(d), 2(c) |
| (h) | (0.000) 300 kg | 3 | Rule 2(d), 2(c) |
| (i) | $6.70 \times 10^4$ mm | 3 | Rule 2(c) |
| (j) | $4.03 \times 10^5$ t | 3 | Rule 2(a) |
| (k) | $6.000 \times 10^{-2}$ m$^3$ | 4 | Rule 2(c) |
| (l) | (0.0)47 cm | 2 | Rule 2(d) |

11.

(a) 34.0 g
(b) 11.1 L
(c) 189.3 m
(d) 0.05 kg
(e) 20 g
(f) 4.5 J
(g) 1.05 mol
(h) 110 cm$^3$
(i) 4020 mg

## Exercise 2

1. Using the information given in this lesson, give the metric equivalents for the following:

(a) 454 g = _____ kg

(b) 50 g = _____ mg

(c) 3 L = _____ $cm^3$

(d) 16 mL = _____ L

(e) 15 mm = _____ cm

(f) 150 $cm^3$ = _____ L

(g) 2.5 kg = _____ g

(h) 4.5 L = _____ mL

(i) 50 km = _____ m

(j) 76 m = _____ cm

(k) 675 mm = _____ m

(l) 0.123 m = _____ mm

(m) 42 m = _____ km

2. The dimensions of a "Black Box", which will be used later in an experiment, were recorded as follows: length 16 cm, width 10 cm, height 80 mm.

Calculate the volume of this box.

3. (a) A dollar bill is about _____ cm x _____ cm.

(b) The average weight of a newborn baby is about: (circle)

(i) 300 kg (ii) 30 kg (iii) 3 kg

(c) The thickness of a dime would be about

(i) 5 mm (ii) 1 mm (iii) 0.1 mm

4. Calculate the volume of 47.25 g of lead.

5. Determine the mass of 14.1 $cm^3$ of iron.

6. A rectangular piece of metal has dimensions 5.50 cm x 4.20 cm x 6.80 cm and mass 3.032 kg. Determine the density of the metal and state the name of the metal.

7. A student was determining the density of an object. He found the mass to be 6.75 g. When the object was immersed in a graduated cylinder of water, the water level rose from 5.3 $cm^3$ to 8.7 $cm^3$. Determine the density of the object.

## Exercise 3

1. State the number of significant digits in each:

(a) 37.3 g _____ (f) $3.9 \times 10^2$ cm _____

(b) 1500 m _____ (g) $5.3 \times 10^{-3}$ mL _____

(c) 12.0 km _____ (h) $3.050 \times 10^2$ m _____

(d) 0.000 2 kg _____ (i) 100 cm/m _____

(e) 0.017 0 L _____ (j) 100 cm _____

2. Perform the calculations and round off to the correct number of significant figures.

(a)
$$\begin{array}{r} 1.28 \\ 0.49 \\ +\,0.062 \\ \hline \end{array}$$

(b)
$$\begin{array}{r} 0.001 \\ 0.082 \\ +\,0.03 \\ \hline \end{array}$$

(c)
$$\begin{array}{r} 6.4 \times 10^3 \\ .275 \times 10^3 \\ +\,25.95 \times 10^3 \\ \hline \end{array}$$

(d)
$$\begin{array}{r} 95.3 \\ -\,6.15 \\ \hline \end{array}$$

(e)
$$\begin{array}{r} 3.006 \\ -1.4 \\ \hline \end{array}$$

(f)
$$\begin{array}{r} 43.43 \\ -\,0.002\ 9 \\ \hline \end{array}$$

(g) $12.9 \times 1.23$ = _____

(h) $2.72 \times 3.1$ = _____

(i) $71.45 \div 0.99$ = _____

(j) $9 \div 21.4$ = _____

(k) $12 \times 15.86$ = _____

## Exercise 4

1. Write in scientific notation:

(a) $448 \times 10^4$ _____

(b) $332 \times 10^{-5}$ _____

(c) $0.022\ 4 \times 10^3$ _____

(d) $0.007\ 5 \times 10^{-3}$ _____

(e) 5000 _____

(f) 0.000 082 _____

2. A student was given the mass of three samples of a compound.

Sample A: $2.5 \times 10^2$ g
Sample B: 300 g
Sample C: $5.3 \times 10^3$ g

Calculate the total mass of the samples and write your final answer in scientific notation.

3. Perform the calculations (write final answers in scientific notation)

(a) $(3.2 \times 10^2)(6.7 \times 10^3)$ _____

(b) $(2.5 \times 10^7)(4.3 \times 10^{-4})$ _____

(c) $\dfrac{2.7 \times 10^8}{3 \times 10^5}$ _____

(d) $\dfrac{1.56 \times 10^{-8}}{1.75 \times 10^{-10}}$ _____

**End of Lesson A**

**COMPLETE THIS FORM AND RETURN WITH LESSON ONE**

Chemistry 30A Course Code 5012

Name _____ Age ☐☐ File Number ☐☐☐☐☐☐☐

Address _____ Phone Number ☐☐☐☐☐☐☐

_____ Postal Code ☐☐☐ ☐☐☐

1. Name of present school (if in attendance). _____

2. Give your final mark (if any) in the following subjects:

| | | | | |
|---|---|---|---|---|
| Science 9 ___ | Chemistry 10 ___ | Mathematics 10 ___ | Biology 10 ___ | Physics 10 ___ |
| Mathematics 9 ___ | Chemistry 20 ___ | Mathematics 20 ___ | Biology 20 ___ | Physics 20 ___ |
| | Chemistry 30 ___ | Mathematics 30 ___ | Biology 30 ___ | Physics 30 ___ |

3. Do you plan to purchase a laboratory kit? _____

4. List any special qualifications or handicaps (jobs, illness, disabilities, etc.) which may influence your progress in this course. _____

5. What is your reason for enrolling in this course? _____

6. List any other subjects you are studying by correspondence instruction. _____

| Lesson | In | Out | Grade | | Lesson | In | Out | Grade |
|---|---|---|---|---|---|---|---|---|
| 1 | | | | | A | | | |
| 2 | | | | | B | | | |
| 3 | | | | | C | | | |
| 4 | | | | | D | | | |
| 5 | | | | | E | | | |
| 6 | | | | | F | | | |
| 7 | | | | | G | | | |
| 8 | | | | | H | | | |
| 9 | | | | | I | | | |
| 10 | | | | | J | | | |
| 11 | | | | | | | | |
| 12 | | | | | | | | |
| 13 | | | | | | | | |
| 14 | | | | | | | | |
| 15 | | | | | | | | |
| 16 | | | | | | | | |
| 17 | | | | | | | | |
| 18 | | | | | | | | |
| 19 | | | | | | | | |
| 20 | | | | | | | | |
| EXAM | | | | | | | | |

**Lesson Grade** ☐ **Exam Portion** ☐

**Lesson Portion** ☐

**Final Grade** ☐

# Chemistry Upgrading
# Lesson B

**DO NOT** send this lesson for correction

# CLASSIFICATION OF MATTER

A. Physical and Chemical Properties
B. Physical and Chemical Changes
C. Classification of Matter
    1. Homogeneous or Heterogeneous
    2. Metals or Nonmetals
    3. Mixtures or Pure Substances
    4. Ionic or Molecular Substances
D. Mixtures, Pure Substances, Elements and Compounds
E. Energy and Chemical Changes
F. Symbols
G. Ways of Separating Matter

## A. Physical and Chemical Properties

Physical properties can be investigated without changing the chemical composition of a substance. If you take any material and do anything with it, without involving any other chemical, you are measuring or determining physical properties. Density (or heaviness), hardness, melting point and other such qualities are examples of physical properties since no other chemicals are involved.

On the other hand, when describing chemical properties, you must involve at least two substances and indicate how they are capable of reacting with one another. For instance, the gases at the far right hand side of the periodic table are inert which means that they are not capable of reacting with other substances. Substances which are capable of reacting with other substances are active. How readily one substance interacts with another substance to produce a third substance is a measure of its chemical properties.

Sometimes, there are gray areas where it is not readily apparent whether a property is physical or chemical. For example, if you mix oxygen and nitrogen, you may end up with a physical diffusion of one gas into the other. Gases can mix together, liquids can mix together, and solids, can physically diffuse into a liquid, without a change in the chemical composition of the substance. It is a physical property of sugar, for example, that it will dissolve in water, but it won't "react" with water in the sense that the sugar molecules or water molecules don't change their structure. On the other hand, lithium will not dissolve in water, but it will chemically react with water to produce hydrogen gas.

Exercise 1

Classify the following properties as either physical or chemical (P or C).

(a) Water boils at 100°C. _____

(b) Argon does not readily react with other substances. _____

(c) Iodine can dissolve in carbon tetrachloride. _____

(d) Iodine does not dissolve in water. _____

(e) Sodium can react with water to produce hydrogen. _____

(f) Gold can be pounded into extremely thin sheets. _____

(g) Silver tarnishes easily at room temperature. _____

## B. Physical and Chemical Changes

When a piece of paper is cut in half, this is a physical change since no new material is formed. Phase changes are also physical changes. A phase change occurs when a solid melts, when a liquid evaporates, when a vapour condenses or when a liquid freezes. At room temperature, as you probably know, oxygen and hydrogen are gases while water is a liquid. Water ($H_2O$) is composed of hydrogen and oxygen. What happens when water evaporates from a dish? (a) Do the water molecules turn into hydrogen and oxygen gas, or (b) do the water molecules remain as water molecules, but in effect dissolved in air just like sugar dissolves in tea ? (a) or (b)? _____

You probably know that oxygen makes everything burn brighter and faster which is the reason that no smoking is allowed in hospital rooms where oxygen is being used, whereas water dampens fires. Would something burn brighter or not so brightly if it were very foggy outside? What would you call a phase change: a chemical change or a physical change? _____

Hint: If water evaporates, do you get more water or more oxygen in the air?

When ice is heated, water results. When water is cooled, ice results. When white bread is put into a toaster, toast results, but what happens when toast is put into a fridge? Do you get white bread back? Obviously, a chemical change occurred when bread was toasted but only a physical change occurred when ice melted.

A chemical change occurs when one or more substances react to produce a new substance which has entirely different chemical and physical properties. A physical change occurs when a single substance is deformed in some way.

Exercise 2

Classify the following changes as physical or chemical changes. (P or C)

(a) boiling of water at 100°C _____

(b) evaporation of water at 20°C _____

(c) dissolving sugar in water _____

(d) exploding dynamite _____

(e) completely blending several pieces of fruits and vegetables in a blender _____

(f) the digestion of the completely blended food _____

(g) the rusting of iron _____

(h) the burning of gasoline _____

(i) a sample of ocean water is evaporated and a white solid residue is obtained _____

(j) milk, on a hot day, turns sour _____

(k) hydrogen and oxygen are put into one container but no different substances are produced _____

(l) hydrogen and oxygen are ignited by a spark to produce water vapour _____

(m) heating dough to make cake _____

hint: will cooling the cake afterwards produce the dough again? _____

## C. Classification of Matter

1. Homogeneous or Heterogeneous
2. Metals or Nonmetals
3. Mixtures or Pure Substances
4. Ionic or Molecular Substances

As can be seen from the above classifications, a quantity of matter can be described in various ways. Also, a given quantity of matter can fit several of these classifications at the same time. For example, you could have a heterogeneous mixture of nonmetals. An example of this would be a bag filled with a book, a plastic pen, and some sulfur.

Homogeneous substances look alike throughout. For example, a mixture of gases in a bottle or a completely dissolved component in a liquid results in a homogeneous solution. When you add sugar to tea and do not stir it, the sugar will settle to the bottom. This would be a heterogeneous solution. When the sugar is completely stirred, the solution is homogeneous.

You are all familiar with metals like iron or steel or copper, and you are familiar with nonmetals like paper and glass. Generally, metals are hard, and have a high luster, and are malleable and ductile, and are good conductors of heat and electricity. Nonmetals generally don't have these properties.

However, some physical properties of metals and nonmetals may be the same. For example mercury, a metal, and bromine, a nonmetal, are both liquids at room temperature. You might ask, what makes a metal a metal? The distinguishing feature of all metals is that they all have similar chemical properties, even though physical properties may greatly differ. All metals readily lose their outer electrons in chemical reactions whereas nonmetals gain electrons in chemical reactions. (Electrons will be covered in greater detail later).

Mixtures and pure substances are different in that a mixture will always consist of several pure substances mixed together, either homogeneously or heterogeneously. Water is a pure substance as well as sugar but when the two are mixed, a mixture forms even though you may not be able to tell the difference between pure water and water with sugar dissolved in it. The components of a mixture cannot be chemically combined.

To discuss the difference between ionic and molecular substances may be premature since electrons have not been discussed yet. For now, all you need to know is that electrons are negative charges and protons are positive charges. If a substance has an equal number of electrons and protons it is neutral and is therefore a molecular substance. An ionic substance has either lost electrons or it has gained electrons. If a substance, such as a metal, loses electrons, it will have more protons than electrons and it will be positively charged. If a nonmetal gains electrons, resulting in more electrons than protons, you will have an ion with a negative charge.

Exercise 3: Classification of Matter

1. Classify the following as either mixtures or pure substances: sugar, salt, a watermelon, a shovel full of dirt, the contents of a pencil sharpener, vinegar, distilled water, mercury, milk, and a flask filled with air.

| Mixtures | Pure Substances |
|---|---|
| | |
| | |
| | |
| | |
| | |
| | |
| | |

2. Classify the following as either homogeneous or heterogeneous: (A) sugar and water well stirred, (B) well stirred equal volumes of oil and water, (C) well stirred mixture of mercury and water, (D) a mixture of ice and water, (E) vinegar, (F) a well mixed mixture of marbles and salt, (G) a 0.0001 $mol/m^3$ solution of HCl, (H) distilled water, (I) a shovel full of dirt. (Use appropriate letters.)

| homogeneous | heterogeneous |
|---|---|
| | |
| | |
| | |
| | |
| | |
| | |

3. List five substances, or parts of substances in your home, which are metallic.

(a) _____

(b) _____

(c) _____

(d) _____

(e) _____

4. List five substances, or parts of substances in your home, which are non metallic.

(a) _____

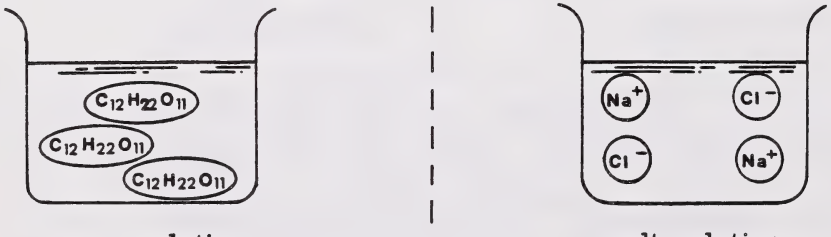
(b) _____

(c) _____

(d) _____

(e) _____

5. The following diagrams indicate how sugar ($C_{12}H_{22}O_{11}$) and how salt (NaCl) dissolves in water.




sugar solution salt solution

(a) Which is an ionic solution? _____

(b) Which is a molecular solution? _____

(You may have to look up the definition of ion in a dictionary)

D. Mixtures, Pure Substances, Elements and Compounds

Although this section overlaps the previous section, the classification of matter will be dealt with a bit further. It was seen in the previous section that a pure substance has the same single substance throughout. You have 100% of one substance and nothing else in a pure substance. In a mixture one has at least two different substances present.

Pure substances can be further divided into elements and compounds. A compound contains two or more elements chemically combined in a definite proportion by weight, which is basically what the Law of Definite Proportions says. A compound can be broken down into its elements. When a compound is broken down into its elements, what kind of a change would this be, a physical change or a chemical change? _____. An element is a substance which cannot be broken down into simpler substances by chemical means.

When looking at pure substances, it is not easy to tell if you have an element or a compound without doing chemical tests. Since only two elements are liquid at room temperature, bromine (red) and mercury (silver), it is a safe bet that any other pure liquid you will come across at room temperature is a compound. Also, unless you have something made of a pure metal, such as copper wires or aluminum foil, it is safe to assume it is an alloy of two or more metals or a compound of some sort. All organic substances such as plastics, paper, wood, etc. are compounds of carbon and hydrogen with possibly oxygen and other elements. The name of the substance might also give it away as to whether something is an element or a compound. For instance, the coating on a tin can is "tin" which is one of the elements, but the coating on a fence is paint and "paint" is not listed as one of the elements. Of course, you will always find exceptions. Can you think of a form of pure carbon which is called by some other very familiar name? _____

As contrasted with compounds, mixtures can be separated or produced by physical processes. Also, with mixtures, there is no definite proportion that the substances have to be in.

Exercise 4

1. Divide the following pure substances into elements and compounds: water, gasoline, magnesium ribbon, paper, pure liquid oxygen, pure liquid acetic acid, zinc, glass, carbon monoxide, nickel and hydrogen gas.

| elements | compounds |
|---|---|
| | |
| | |
| | |
| | |
| | |
| | |

2. Define an element. _____

_____

_____

3. Define a compound. _____

_____

_____

4. Since uranium is radioactive and is capable of converting itself to lead, would you classify uranium as an element or a compound?

_____

Why? _____

_____

_____

5. How could you show that a salt water solution is a mixture and not a compound?

_____

_____

_____

_____

_____

6. Mixtures can be separated into compounds or elements by simple physical or mechanical means. True or False _____

7. Compounds can be separated into elements by more complex physical or mechanical means. True or False _____

8. Pure acetic acid has 40% carbon, 6.7% hydrogen and 53.3% oxygen. If you had a substance which contains 30% carbon, 9% hydrogen and 61% oxygen, could it still be pure acetic acid? _____ Why?

_____

_____

9. Milk has 87% water with minute amounts of minerals, fats, proteins and other substances. Would you still have milk if you had 92% water and 8% of the other substances? _____ Why? _____

_____

_____

10. Different elements have greatly different properties. It is important to use each element for its beneficial properties, but to be careful that any given element does not harm us. For instance, tungsten has an extremely high melting point, and when alloyed with steel, gives the steel great strength at high temperatures. Chromium, which is very corrosion resistant, is added to iron to produce stainless steel. Other metals, such as mercury and lead, have many good uses, but can also be

harmful or poisonous to man. Which metal is commonly used for wiring a house? _____ What is a very light pot likely to made of? _____ Which metal is often the filament of a bulb? _____ A galvanized pail is made of iron, and has a thin coating of _____. Which element is put into balloons to make them float? _____ Which element has to be put into a tank to enable sea divers to breath? _____ Which element would be an important source of nuclear energy in power plants? _____ Seventy percent of the sun is composed of _____.

11. State the Law of Definite Proportions. _____

_____

_____

12. Tin forms two oxides. One contains 78.77% tin and the other 88.12% tin. Show that these oxides illustrate the Law of Multiple Proportions. The Law of Multiple Proportions states that if a fixed weight of one element is chosen, you must find the weight of the other element in two different compounds. In the second case, the weight will be an exact multiple of 2 or 3 etc. as compared with the first case. (Hint: arbitrarily assume you have 100 g of each compound). In a 100 g sample of Compound A you would have 78.77 g of tin and _____ g of oxygen. Therefore one gram of tin combines with _____ g of oxygen. On the other hand, Compound B would have 88.12 g of tin and _____ g of oxygen. Therefore one gram of tin combines with _____ g of oxygen.

For a fixed mass of tin, what is the mass ratio of oxygen in the two compounds? _____

Does this illustrate the Law of Multiple Proportions? _____

## E. Energy and Chemical Changes

As you learned in the previous lesson, energy cannot be created nor destroyed. When chemical reactions occur, some bonds are broken, and other bonds are formed. The new bonds may be either weaker or stronger than the old bonds. Since you cannot get anything for nothing, however, the reaction may either produce heat, or it may absorb heat.

As you will learn in greater detail later, all elements are composed of negatively charged electrons and positively charged protons. Analogous to magnets, like charges repel, and unlike charges attract. In every compound, there is an optimum distance for the various elements to be separated, since various attractive and repulsive forces are at work. Basically, a bond forms because electrons of one element are attracted to the protons of the next element. Due to repulsive forces between two different nuclei, and between electrons from two elements, the elements in a compound can't get too close together. Since both attractive and repulsive forces are acting in a compound, the situation is more complex than that of two magnets.

### Exercise 5

1. When a chemical reaction occurs, the products may have less potential energy than the reactants. Since total energy is not lost in a chemical change, would you expect the products to be (a) hotter or (b) colder than the reactants? _____.

2. When you have ice at 0°C, and you add just enough heat to convert everything to water at 0°C, all particles have the same average speed or kinetic energy, since the temperature is the same. Was this added heat wasted? _____. Is it fair to assume that the added heat did work in separating the various ice molecules from each other? _____.

3. In the above case, heat was added to make a phase change occur. Is it reasonable to assume that at room temperature, certain chemicals do not have enough energy to react, but added heat may cause a chemical reaction to occur? _____.

## F. Symbols

Berzelius devised the symbol system that we use today. In this system, every element is represented by one or two letters. The first letter is always capitalized. If there is a second letter, it is always written in the lower case. Often, the first, or first two letters are the letters in the element's name when it is pronounced in English. For example, the symbol

for carbon is C, and for calcium it is Ca. The attached periodic table gives the symbols for all elements. There is no need to memorize all symbols and elements, although you will probably automatically remember the names and symbols for some of the 106 elements which are known today. Spend a few minutes studying all the information that is given on the table. At this stage, much of what is listed may not mean much to you, but hopefully, more things will mean something to you by the time you have finished Lesson J. Keep this periodic table handy as you do your future lessons.

Exercise 6

1. Following, are the nine most abundant elements in the earth's lithosphere and atmosphere. Beside each element, give its chemical symbol.

| | | | |
|---|---|---|---|
| oxygen | _____ | hydrogen | _____ |
| calcium | _____ | aluminium | _____ |
| magnesium | _____ | potassium | _____ |
| silicon | _____ | iron | _____ |
| sodium | _____ | | |

2. Which of these elements are not one of the first twenty elements of the periodic table? _____.

3. Besides the first twenty elements, other elements you will often come across are the following. Beside each element, give its symbol.

| | | | | | |
|---|---|---|---|---|---|
| antimony | _____ | arsenic | _____ | barium | _____ |
| bromine | _____ | chromium | _____ | cobalt | _____ |
| copper | _____ | gold | _____ | iodine | _____ |
| manganese | _____ | mercury | _____ | lead | _____ |
| platinum | _____ | silver | _____ | nickel | _____ |
| uranium | _____ | zinc | _____ | tin | _____ |

## G. Ways of Separating Matter

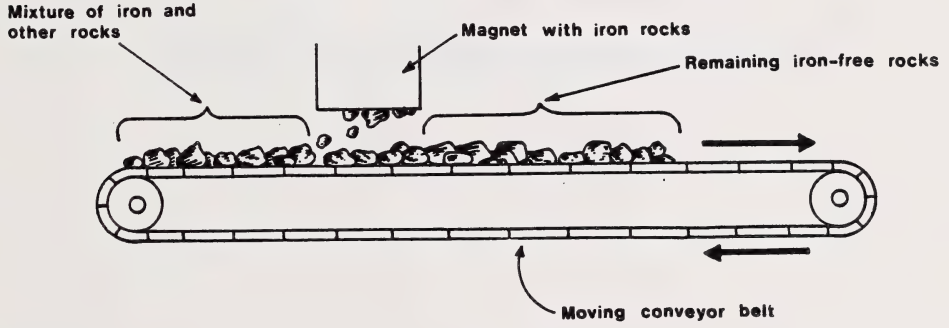
As you learned before, mixtures can be separated by physical means since the various components are not combined chemically. Compounds, however, can be separated into its elements by chemical means.

Various means can be used to separate mixtures depending on their different properties. For example, some differences in physical properties are color, odor, taste, freezing point, magnetic properties, solubility in water, density, hardness, malleability, ductility, luster, electrical conductivity and others. (Always have a dictionary handy for words you don't know the meanings of.)

You can use your own imagination as to how you can separate substances using these different physical properties. Ways in which malleability and magnetism can be used will be discussed below.

Silver is very malleable which means that it can be pounded into thin sheets. If we were to give you twenty small pieces of "rock", about the size of marbles, and told you that eight are silver and the other twelve are just ordinary rock, — assuming that color, density, the outer texture of the rocks and all other physical properties are the same, except for the malleability, how would you separate the mixture? The simplest way would be to hit each rock with a hammer. The silver will be deformed like a piece of putty but the other rocks will tend to splinter up and crack.

Iron, nickel and cobalt along with many compounds containing these and other elements can be separated from other rocks by using their ability to be attracted to magnets. Below, is an illustration as to how a copper mill can separate the iron from its crushed rocks.




Exercise 7

1. You are given two "silver" dollars, one made of pure silver and the other one made of plastic but coated with a thin layer of silver colored paint so it looked exactly like the silver dollar in every way. Give two physical tests you can use to determine which is which, without damaging any piece in the smallest way. Assume you have all necessary equipment.

   (a) _____

   (b) _____

2. If you had a mixture of two types of crushed rocks, which were insoluble in water, how would you separate them if one type of rock had a density of 1.2 $g/cm^3$ and the other type had a density of 0.85 $g/cm^3$ ?

3. You have a mixture of sand and salt which is well mixed, and the particles look alike. Assuming that the sand would just sink to the bottom in water, and that the salt completely dissolves, give the steps you might go through to end up with a separated pile of each substance.

END OF LESSON B

# Chemistry Upgrading
# Lesson C

**DO NOT** send this lesson for correction

CONTENTS

A. Experiment 1 – Black Box Experiment
B. Development of an Atomic Model
C. Rutherford's Experiments
D. Periodic Law
E. Atomic Theory
F. Gay Lussac's Law of Volumes
G. Avogadro's Principle.

## A. Experiment 1 – Black Box Experiment

Scientists cannot see inside the atom directly, so much of what they know about the atom will have to come from indirect evidence. The purpose of this experiment is to have you try to make deductions about objects from indirect observations. Chances are that you have probably seen the objects you are trying to identify at one time or another. However, scientists have to try to figure out the identity of objects which they have never seen before.

Procedure:

Have someone else put 3 small unbreakable objects inside a box. This box should be about the size of a shoe box. The box must then be taped shut, and you are not to see the objects beforehand, nor should you be able to see inside the box. Now, you are ready to try to determine the contents of the box.

Any manipulations which do not involve opening the box and do not damage the box or its contents are permissible.

Fill in the table below, using the hypothesis you are given and add a few of your own. Since a hypothesis is something you can test, test your hypothesis and give an "observation" of what you see, hear, feel or smell. As well, give an interpretation based on your observation. One is given to get you started.

| Hypothesis: | Observation: (for each object) | Interpretation: (for each object) |
|---|---|---|
| 1. The box contains 2 objects | 3 distinct sounds are heard when the objects fall to an end. | The box has at least 3 objects. |
| 2. The objects are completely round. | (a) | (a) |
| | (b) | (b) |
| | (c) | (c) |

| Hypothesis: | Observation: | Interpretation: |
|---|---|---|
| 3. All objects are the same. | | |
| 4. All objects are magnetic. (test with a magnet) | | |
| 5. The objects are very large. | | |
| 6. The objects are very heavy. | | |
| 7. | | |
| 8. | | |

Exercise 1

1. Based on your interpretation as to what the various objects are like, what is your conclusion about the three objects. If you are not sure what the objects are, at least draw a diagram as to what you think each object approximately looks like.

2. Why is this called a black box experiment?

3. Would you call cures for cancer or the internal structure of the atom as present day black boxes? Explain.

It must be frustrating to try to guess the contents of this box when it would be so simple to open it. However, the whole idea of the experiment was to indicate that scientists often have to go through similar experiences because one can't look into something that easily and come up with a solution, for example a cure for cancer. Hopefully, you realized that much information could be gathered by indirect methods and with a little imagination, you can draw pictures of what you think was in the box. You have probably seen the contents of the box at one time or another. A scientist, however, who is trying to describe an atom, may have a real probelm as he is trying to describe something which he has never seen, but from which he has much indirect evidence. One real problem which has not been solved yet is to make a model which adequately describes the complete nature of light. Some experiments indicate that light is made of waves, while other experiments indicate that light is made of particles. At this stage we do not seem to be intelligent enough to come up with a model to explain both observed properties at the same time.

Now open the box and check your conclusions. You can see the contents. Can you think of any other tests that you could have used to give you a more accurate model if you had the available equipment?

_____

_____

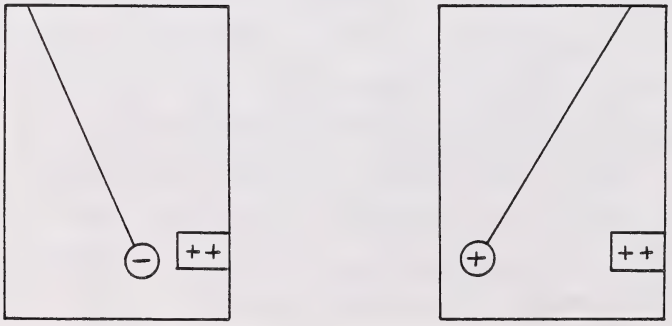
B. Development of an Atomic Model

In the following sections, we will see how scientists had to use similar indirect methods, that you used in the Black Box Experiment, to try to come up with a model of the atom.

Although scientists talked about the atom as the essential particle of matter, a suitable atomic model was difficult to develop.

The subject of electricity furnished the first clue to the puzzle about the nature of the atom. Early scientists studied static electricity, which is the kind that discharges from your fingertips when you touch a metal doorknob or radiator after shuffling across the rug.

Elementary science courses today show how a glass rod rubbed with silk, or an amber rod rubbed with wool will pick up bits of material the way a magnet picks up pieces of iron. You may try a series of simple and entertaining experiments with a small pith ball on a thread which will show that apparently there are two kinds of electricity. Matter containing kind A is repelled by matter that also contains kind A. Anything containing kind B will attract anything containing kind A. The two kinds of electricity are positive and negative. Positively charged particles repel all other positively charged particles; negatively charged particles repel other negatively charged ones. Positively charged particles attract negatively charged particles. Neutral particles are those having no charge.



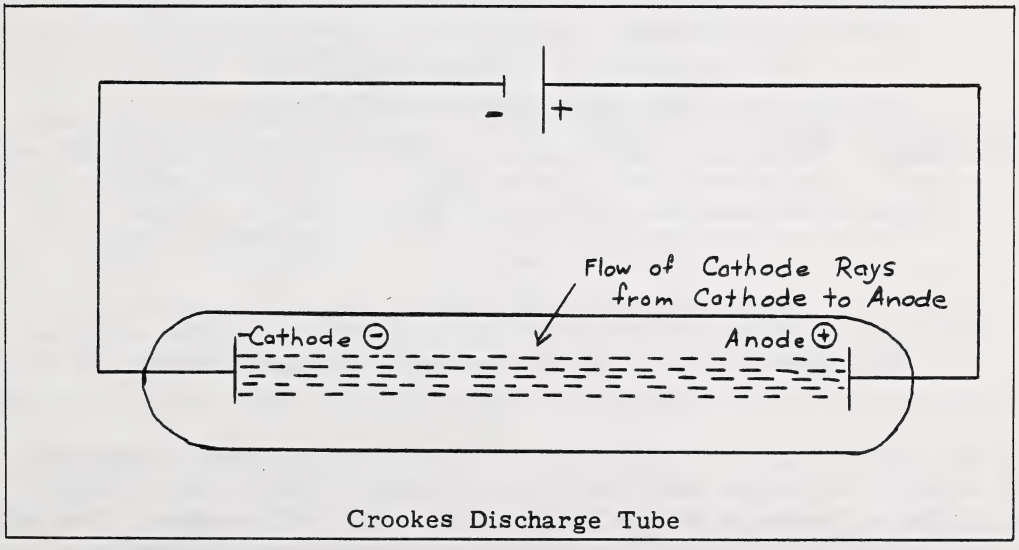

Electrostatic Attraction and Repulsion

## Evidence for Electrons

Among the many kinds of electrical phenomena that fascinated physicists, who invented many devices to illustrate the strange properties of electricity, was the spark that leaped through the air between two highly charged but oppositely charged electrodes.

This discharge of electricity through a gas produced certain observable reactions in the gas itself; for instance electrical discharge produced different colours in different gases. Even more remarkable differences were obtained by sending discharges through a partial vacuum in a glass tube containing negative and positive electrodes sealed into its opposite ends. This tube could also be filled with pure gases at different pressures with different results. Sir William Crookes is credited with refinements in the making of this apparatus, which is called the Crookes tube.

When a high-voltage electrical current is applied to the ends of a Crookes tube filled with any kind of gas, the electrical charge which by convention is called negative builds up in the tube's negative electrode – the cathode. As electricity is fed into the tube the charges continue to build up on the cathode and when sufficiently large they bolt across to the anode. If the gas is pumped out of the tube, instead of an intermittent series of spark discharges, a steady greenish glow appears on the tube wall opposite the cathode indicating that something coming from the cathode hits the glass and causes the glow. These "radiations" were named cathode rays.



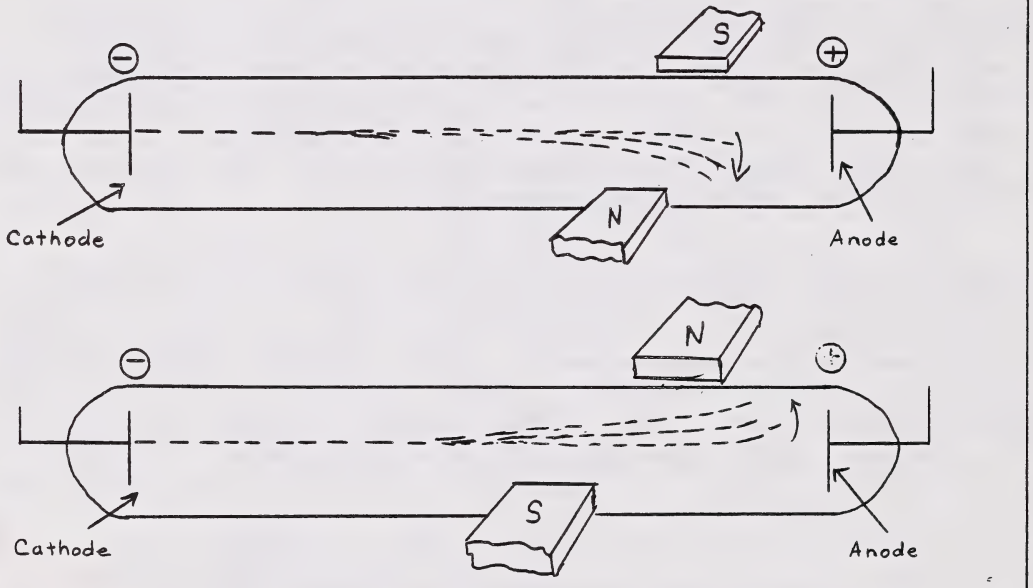

Crookes Discharge Tube

Further investigations yielded the following conclusions regarding the nature of cathode rays.

An object placed in the path of the rays cast a shadow on the fluorescing wall of the tube — a shadow that was a faithful reproduction of the object. This could mean only that cathode rays travelled in straight lines from the cathode.

Next, a little paddle wheel, placed on rails that ran the entire length of the tube, was pushed away from the cathode when the radiation hit the paddles. This proved that the rays must consist of solid particles, or the paddle wheel would not have moved.

Finally, a magnet held close to the tube bent the cathode rays from their straight path, proving that they consist of electrically charged particles. Hence, the rays are also deflected by electric charges, and the direction of the deflection shows that the charge carried by the cathode-ray particles must be negative and not positive.




The Deflection of Cathode Rays in a Magnetic Field

It was legitimate to conclude that cathode rays consist of extremely small particles of matter that have mass, volume and a negative charge. These particles were named electrons. The amount of electricity that each electron carries was considered to be the smallest possible quantity of electricity that could exist by itself.

Since the rays stopped when the current of electricity was shut off, these particles must have had to come from the material of the cathode. Thus electrons must be part of matter itself and related to matter in such a way that normally they cannot be detected in neutral objects.

## Evidence for Protons

Every atom and therefore every molecule which consists of atoms must contain electrons that can be pried out and sent speeding through space. The remaining bulk of the atom, robbed of some of its electrons, must be positively charged. Such positively charged atoms or molecules would be attracted to the cathode (negative electrode) or any negatively charged object.

Following this reasoning, the Crookes tube was searched for positively charged particles and they were found – a thin trickle of them moving away from the anode (+) to the cathode (-). Various experiments left no doubt that if the tube contained a little bit of gas, the streaming electrons of the cathode "ray" hit the gas molecules and knocked a few electrons out of them. The gas molecules were now left positively charged and moved towards the cathode.

## Radioactivity

Clues to the complexity of the atom came from another field altogether as a result of the discovery of radioactivity. By the turn of the twentieth century it had been established that uranium and several other newly discovered very heavy elements spontaneously emitted three kinds of radiation, which were named alpha, beta, and gamma rays.

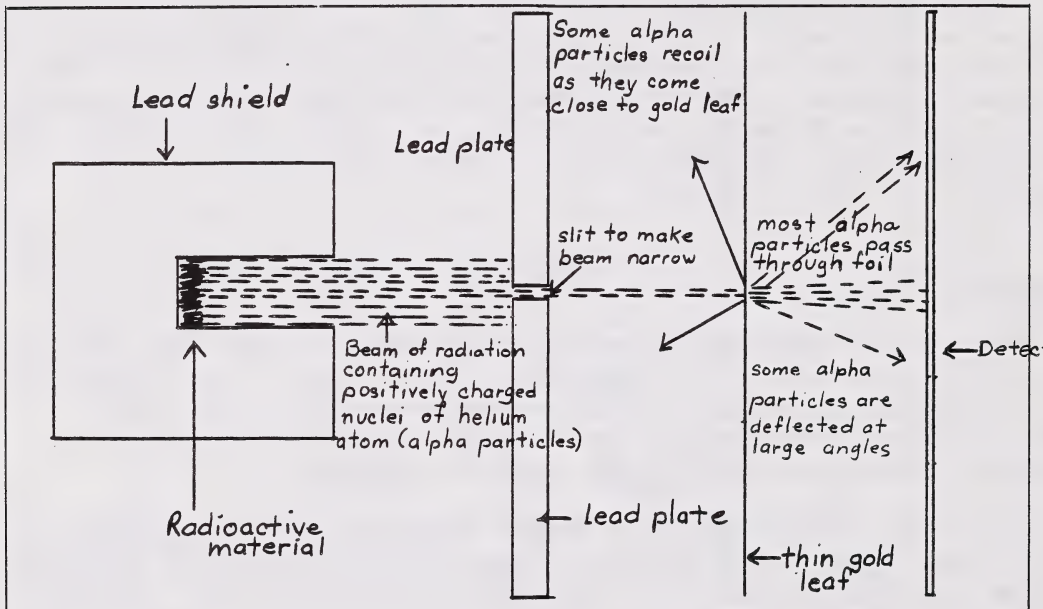
Alpha rays ( $\alpha$ ) consist of a stream of positively charged particles with a mass four times that of hydrogen. Beta rays ( $\beta$ ) consist of negatively charged particles identical to the electrons of the cathode rays. Gamma rays ( $\gamma$ ) do not consist of particles at all but are electromagnetic waves like light but of a shorter wave length.

Plainly, both negative and positive particles must exist in atoms and can be found there without benefit of electrical currents. This was absolute proof of the electrical nature of atoms.

## C. Rutherford's Experiment

Several models of the atom were proposed to illustrate what was known, but let us consider only the one that has become most useful for modern scientists. The British physicist Rutherford set up the experiment that was to prove the basis for the modern atomic model.

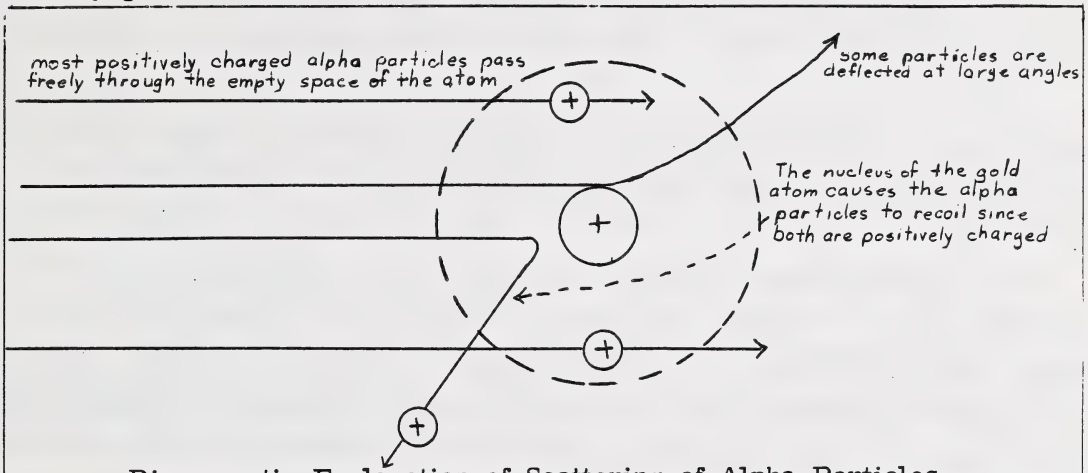
He bombarded a target of very thin gold leaf with alpha particles emitted by a radioactive element. These alpha particles as already mentioned are four times as heavy as hydrogen and are in reality positively charged residues of helium atoms. They travel at a speed of roughly 16 000 kilometres per second. Most of these particles went right through the gold leaf in a direct line to hit the detection screen behind the gold leaf – as though the gold atoms of the leaf were empty space. Some of the particles, however, were deflected to the side and hit the screen at different distances from the screen's center. Furthermore a few of the particles were knocked completely backward without ever going through the gold leaf.


Some alpha particles recoil as they come close to gold leaf
Lead shield
Lead plate
slit to make beam narrow
most alpha particles pass through foil
Detector
Beam of radiation containing positively charged nuclei of helium atom (alpha particles)
some alpha particles are deflected at large angles
Radioactive material
Lead plate
thin gold leaf

Rutherford's Experiment on the Scattering of Alpha Particles.

Rutherford's explanation was the following. Atoms of gold were so close together in the solid gold foil that not even light rays could penetrate them. Since positively charged helium atoms did get through the thin foil, they must have gone right through the atoms, not between them, and yet the gold leaf was not damaged. There must be empty space inside the atom, enough empty space to allow an alpha particle free passage through the atom. Some of the alpha particles, however, were knocked aside at different angles. Hence there must be within the atom's total space a very solid compact part. This solid part must have a positive charge on it because several of the positively charged particles were flung right back, more violently repelled than could be accounted for by a direct hit on the solid part.

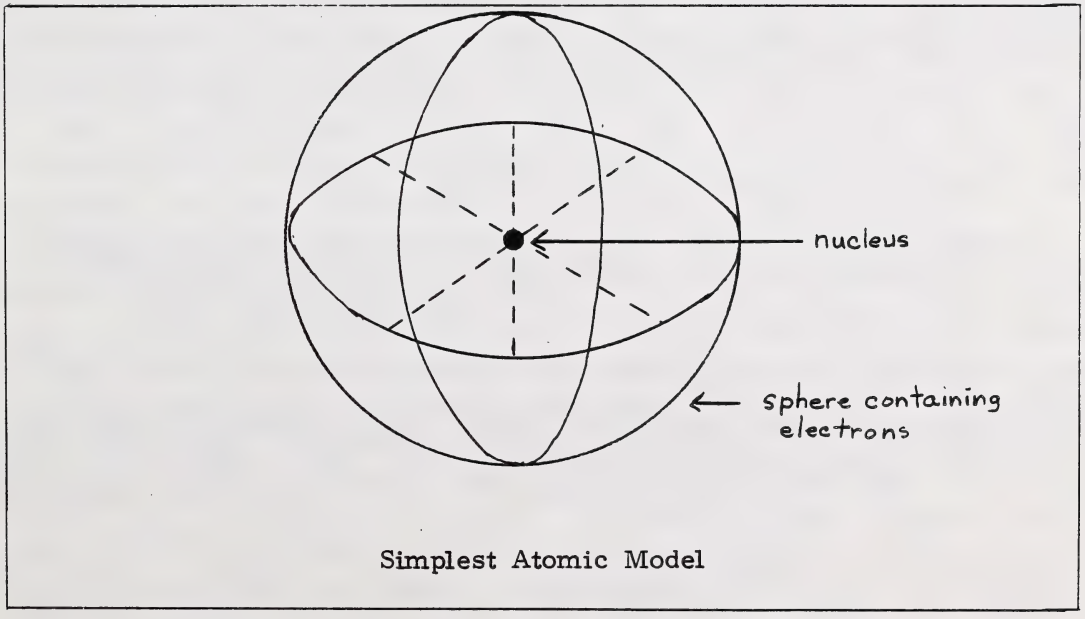
The model Rutherford imagined consists of a center core, and a positively charged nucleus that contains most of the mass of an atom. Since the atom as a whole is neutral, electrons are grouped around the nucleus in such a way that they balance the positive charges. These electrons are at a relatively great distance from the nucleus, creating the surface of the atom.


most positively charged alpha particles pass freely through the empty space of the atom
some particles are deflected at large angles
The nucleus of the gold atom causes the alpha particles to recoil since both are positively charged

Diagramatic Explanation of Scattering of Alpha Particles

## The Atomic Model

The simplest diagram of the atom that we can draw from the above information is that of a sphere with a dot in the center. The dot represents a very small and positively charged nucleus which contains practically all of the mass of the atom. The sphere at a distance represents the space containing the atom's electrons. The mass of an electron has been calculated to be 1/1840 of the mass of a hydrogen atom.




Simplest Atomic Model

Atoms are electrically neutral, and therefore the picture of the simplest atom that we can imagine is one with a nucleus that has a single positive charge balanced by a single negative charge some distance from it. This is, in effect, the hydrogen atom, the smallest and lightest atom. The atoms of all other elements are heavier than hydrogen and therefore have larger nuclei with more than one positive charge and thus have more than one electron at a distance from it.

The idea of atoms with positive centers and negative charges around the center presented a certain problem. Knowing that positive and negative charges attract each other, the obvious question was, "Why does the atom not collapse?"

One way to answer this is to assume that the electron is in motion and therefore will not fall into the center of the atom. You know that it is possible to swing an open pail of water in a circular vertical plane without spilling any of the water. In a similar way, if one considers the electron as being in motion, the motion of the electron will balance the attraction of the positive center. The speed of the electron is close to 100 000 miles a second, which is equivalent to four times around the earth's equator in one second. The positive-negative forces of attraction between electron and nucleus keep the electrons from flying off into space while the centrifugal force of the electrons provided by their own speed keeps them from crashing into the nucleus.

The electrons weave a complex pattern of orbits around the nucleus. These are not all circular but sometimes flatten out into elipses or twist into dumbbell shapes. This again is rather difficult to diagram; however we can calculate with absolute certainty from analysis of crystals in which the nuclear distance is rigidly consistent, the average distance the electron maintains from the ucleus. A solid does not bulge, shrink, or cave-in because of the changing orbits of electrons around the nuclei. In other words at any instant the probability of finding an electron is greatest at a certain distance from the nucleus. The electron at times might be much farther away, but the chances are mathematically such that over a period of time it will be found most often at the stated distance.

Any electron in any atom can be excited by collision with an outside source. Radiations such as X-rays, other fast travelling electrons or another more energetic atom can cause an electron to travel faster, thereby giving it greater energy which allows it to move farther away from the nucleus. But, it moves farther away in leaps and bounds rather than at smooth flow. That is, the electron can occupy any one of several well-defined orbits, or tracks, around the nucleus, depending on how much energy it has; but it cannot revolve between those orbits. An analogy for the energy levels of an atom is the gear mechanism of a car. The car can travel only in first, second, third or fourth gear, never in a fraction of a gear. It was Bohr who reasoned that the electron simply leaps from one groove or orbit into the next if it is sufficiently exicited. It seems to vanish from one level of energy and reappear at a higher level farther from the nucleus if some-

thing has given it the extra energy, or impetus; or it may fall into a level closer to the nucleus if it loses some energy. An excited electron will gradually fall back erratically, perhaps, level by level into its normal channel, the so-called ground level, emitting as it leaps down the energy that it had absorbed when it became excited.

## The Nucleus

After Rutherford's classic experiment on the scattering of alpha particles, further research in which neutral atoms were bombarded by alpha (+) and beta (-) particles shot from radioactive elements eventually led to the construction in the 1930's of the first "atom smasher". From these machines and modern developments of them, we have extracted proof that there are more than thirty bits of basic matter and energy associated with the nucleus. The most familiar of these at the present time are the proton, neutron, neutrino, meson and positron. For each of these particles, and some have several different versions, there is an anti-particle; a collision between a particle and its anti-particle destroys both, giving off energy.

Although many arrangements have been proposed, there is no universally accepted theory as to how all these pieces are assembled into the nucleus. In any case, only the proton, the neutron and the electron have an enduring existence as independent entities. All the other pieces exist only fleetingly, for a small part of a second, then disappear in mysterious ways. The protons and neutrons in the nucleus are collectively called nucleons.

The proton, which was discovered by Rutherford in 1919, has a mass of approximately 1 a.m.u. In chemistry and in much of physics the exact mass of these particles is not used. The proton carries a single positive charge of electricity, equal and opposite to the charge on the electron. Protons are the nuclei of hydrogen atoms.

The neutron is slightly more massive than the proton but in chemical calculations its mass is considered to be 1 a.m.u. The neutron has no electrical charge so it was not discovered until 1932 by Chadwick, many years after the proton was well known. Being neutral, the neutron cannot be handled in electrical and magnetic fields.

The electron has a mass approximately $\frac{1}{1840}$ that of a proton. It has a single negative charge and is thus affected by any electrical or magnetic field in its vicinity. The speed of an electron depends upon the conditions, but it can approach the speed of light and is always in that range of velocity.

For the chemist only these three elementary particles - protons, neutrons and electrons - are important, as they are the only stable constituents of the atom. With these we can explain nearly all the events in a chemical reaction, and all the physical properties of different elements and compounds.

Charge and Mass of Some Fundamental Particles

| Particle | Symbol | Charge | Approximate Mass |
|---|---|---|---|
| Electron | $e^-$ | -1 | 1/1840 a.m.u. |
| Proton | p | +1 | 1 a.m.u. |
| Neutron | n | 0 | 1 a.m.u. |

## The Anatomy of Individual Atoms

To discuss the organization of the individual atoms it is logical to start with hydrogen (H) which consists of 1 proton and 1 electron and has no neutrons. If a proton and an electron from any source are brought together under ordinary conditions they will instantly link together into a hydrogen atom. The mass number of an element is the number of protons plus the number of neutrons.

The next element is helium (He), whose nucleus consists of 2 protons and 2 neutrons and thus has a mass number of 4. Revolving around the nucleus to balance the two positive protons are 2 electrons. Lithium (Li) is the third element, with 3 protons and 4 neutrons in the nucleus and, of course 3 orbiting electrons to balance the positive charges. The mass number of lithium atom is 7. Beryllium (Be), the fourth element, which has (4) protons, (5) neutrons, and 4 electrons, has a mass number of (9) + .

It seems that the elements can be lined up in such a way that each one has exactly one proton more in its nucleus than the previous one, and also one more electron to balance it. The number of neutrons in succeeding elements does not increase in any regular manner. The neutrons do not require electrons to balance them since they are neutral, but they do add mass.

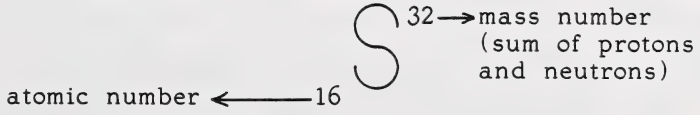
Actually every atom of a particular element has the same number of protons, but there may be a varying number of neutrons associated with the fixed number of protons in the nucleus. All oxygen atoms have 8 protons, since it is the eighth element in the table. But some oxygen atoms have 7 neutrons, some have 8 neutrons and some have 9 neutrons. All atoms of oxygen with 8 protons, but differing in the number of neutrons, are called isotopes of oxygen. All these isotopes exist in nature in a mixture that does not vary in proportion from one place to another and that gives an average mass for the three of 16.

The isotope with a mass number of 16, that is, with 8 neutrons happens to be by far the most common of the three. In other words, if a sample of oxygen taken from the air is analyzed, a very small percentage will consist of atoms containing 7 neutrons, a very small percentage will contain 9 neutrons and the majority will contain 8 neutrons. Their atomic masses or mass number will be 15, 17, and 16. The percentage of each isotope in the mixture is amazingly consistent, no matter where the oxygen is found. Remember that the atoms of each isotope of an element are alike in properties but the different isotopes of an element have different numbers of neutrons. The identifying feature of an element is the number of protons in the atoms, which is the same in all isotopes of the element.

To distinguish between isotopes of the same element the mass number of the element is placed after the name of the element. When carbon-12 was mentioned as the standard for atomic mass, you must have wondered about the significance of the 12. It indicates that the isotope of carbon with a mass number of 12 is the isotope which contains 6 protons and 6 neutrons.

Isotopes can also be represented by symbols. To represent carbon-12 the chemist writes $_{6}C^{12}$. The number in the lower left hand corner indicates the atomic number of the element. The number in the upper right corner indicates the atomic mass number of the isotope.

e.g.




The three isotopes of oxygen can be represented as $_{8}O^{15}$ $_{8}O^{16}$ $_{8}O^{17}$

If the pattern of protons increasing one by one persists for all the 106 elements, then nitrogen (N) the seventh element in our table should have 7 protons and 7 electrons. It does. Krypton (Kr) the thirty-sixth element, should and does have 36 protons. Uranium (U) turns out to have 92 protons in its nucleus together with 92 electrons in its cloud around the nucleus. The most common isotope of uranium contains 146 neutrons - and so the atomic mass or mass number is (92 + 146) or 238. The 92 electrons have a negligible mass. We know that the properties of each element are recognizably and uniquely its own, and quite different from the properties of other elements. Oxygen is nothing like iron, and sodium is nothing like sulphur.

**Terms to Remember**

| | |
|---|---|
| Atomic Number: | The atomic number of an element equals the number of protons in the neutral atom. |
| Mass Number: | The sum of the masses of its protons and neutrons. Except in nuclear research both protons and neutrons are considered to have a mass of 1. Therefore mass number = neutrons + protons. |
| Atomic Mass: | The average mass of an element's isotopes compared to the standard carbon isotope with an atomic mass of 12 a.m.u. |

## D. Periodic Law

When the 106 known elements are arranged in order of increasing atomic numbers (i.e. number of protons in the nucleus) each successive element has different chemical and physical properties from its neighbors. On the other hand, there are striking likenesses among some of the elements. For example only a dozen or so elements are gases at standard temperature and pressure. Of these gases, six are so similar chemically, that instead of arranging the elements in a long horizontal line, we emphasize the reappearance of similar properties by considering them together. They are identified by a family name, the inert or noble gases (lack of reactivity). They consist of helium, neon, argon, krypton, xenon and radon.

The observation that various properties (of which inertness is only one example) recur periodically throughout the entire list of elements is called the periodic law. It was discovered by the Russian scientist, Mendeleev.

Since the recurrence of properties is a regular occurrence, it is convenient to group all of the elements with similar chemical properties into chemical families or groups. If we know about one element in a group we can understand the chemistry of all the other elements in that group. In the periodic table each group appears in a vertical column.

Periodic Table

| Group | I | II | | | | | | | | | | | III | IV | V | VI | VII | O |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| Period 1 | 1 H | | | | | | | | | | | | | | | | | 2 He |
| 2 | 3 Li | 4 Be | | | | | | | | | | | 5 B | 6 C | 7 N | 8 O | 9 F | 10 Ne |
| 3 | 11 Na | 12 Mg | | | | | | | | | | | 13 Aℓ | 14 Si | 15 P | 16 S | 17 Cℓ | 18 Ar |
| 4 | 19 K | 20 Ca | 21 Sc | 22 Ti | 23 V | 24 Cr | 25 Mn | 26 Fe | 27 Co | 28 Ni | 29 Cu | 30 Zn | 31 Ga | 32 Ge | 33 As | 34 Se | 35 Br | 36 Kr |
| 5 | 37 Rb | 38 Sr | 39 Y | 40 Zr | 41 Nb | 42 Mo | 43 Tc | 44 Ru | 45 Rb | 46 Pd | 47 Ag | 48 Cd | 49 In | 50 Sn | 51 Sb | 52 Te | 53 I | 54 Xe |
| 6 | 55 Cs | 56 Ba | 57-71 * | 72 Hf | 73 Ta | 74 W | 75 Re | 76 Os | 77 Ir | 78 Pt | 79 Au | 80 Hg | 81 Tℓ | 82 Pb | 83 Bi | 84 Po | 85 At | 86 Rn |
| 7 | 87 Fr | 88 Ra | 89- + | | | | | | | | | | | | | | | |

| | | | | | | | | | | | | | | | | |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| * | 57 La | 58 Ce | 59 Pr | 60 Na | 61 Pm | 62 Sm | 63 Eu | 64 Gd | 65 Tb | 66 Dy | 67 Ho | 68 Er | 69 Tm | 70 Yb | 71 Lu |
| + | 89 Ac | 90 Th | 91 Pa | 92 U | 93 Np | 94 Pu | 95 Am | 96 Cm | 97 Bk | 98 Cf | 99 Es | 100 Fm | 101 Md | 102 No | 103 Lw |

As you can see by the above table, the basic feature of the periodic table is the arrangement of the elements in order of increasing atomic number with the elements with similar properties placed under each other in a vertical column called a group. There are eight main groups designated in Roman numerals as I, II, III, IV, V, VI, VII and O.

The noble gases in group O are all gases at STP characterized by a lack of reactivity.

The elements that follow the inert gases — lithium (3), sodium (11), potassium (19), rubidium (37), cesium (55) and francium (87) — are all metals. As a group they are called the alkali metals. Their common physical properties are shiny luster and good conductivity of heat and electricity.

The elements that come directly before the inert gases in the table — fluorine (9), chlorine (17), bromine (35), iodine (53), and astatine (85) also resemble one another in physical and chemical properties. As a group they are called the halogens, and are all non-metals (poor conductors of heat and electricity).

The elements that fall between the alkali-metals and the halogens show a progressive gradation of properties between the two extremes. If we look across the periodic table from group I to group VII there is a gradual decrease in metallic properties and an increase in the non-metallic properties of the elements.

The periodic table is very useful in organizing the chemical properties of the elements. We shall consider the periodic occurrence of similar atomic structure in a later lesson.

Exercise 2

1. Complete each of the following statements.

(a) Atoms of the same element containing equal numbers of protons but different numbers of neutrons are called _____.

(b) The number of protons plus the number of neutrons in any element equals the _____.

(c) The number of electrons in an atom equals the number of _____.

(d) The neutral particles in the nucleus of an atom that are approximately the same mass as protons are called _____.

(e) The relative mass of an element compared to that of carbon-12 is an element's _____.

(f) The number of protons in an atom of an element determines its _____.

(g) The beams of electrons visible in Crookes tubes are called _____.

(h) The positively charged densely packed part of an atom is the _____.

(i) The shells surrounding the nucleus contain particles called _____.

(j) Like charges _____. Unlike charges _____.

(k) Radioactive elements may emit three types of rays

(i) _____ (ii) _____ (iii) _____

(l) A neutron has a mass slightly _____ than a proton.

(more)(less)

(m) The positive charge on the nucleus of the atom is due to the presence of _____.

(n) The mass of a proton is approximately equal to _____ times the mass of an electron.

(o) The approximate mass of a proton (and of a neutron) is _____ a.m.u.

2. (a) In the Crookes tube, in what direction does the beam of electrons flow?

_____
(anode to cathode or cathode to anode)

(b) What name was given to the "radiation" in the Crookes tube?

_____

(c) Why did scientists come to the conclusion that cathode rays consist of solid particles?

_____

_____

(d) What caused the cathode rays which normally travelled in straight lines to curve off course?

_____

_____

(e) What conclusion can you arrive at regarding the nature of cathode rays from the evidence of the deflection?

_____

_____

_____

3. Why did Lord Rutherford come to the conclusion that the atom has a nucleus in which the positive charge and the mass are concentrated?

_____

_____

_____

4. Give a short description of the nature of the three types of rays emitted by radioactive elements.

5. In terms of subatomic particles, give an explanation as to why atoms are neutral particles.

6. Since there is attraction between the positively charged nucleus and the electrons in orbits around the nucleus, why does the atom not collapse?

7. Name three subatomic particles that have very brief existences in the nucleus.

8. What are isotopes? Is the atomic mass the same for all isotopes of a given element? Why?

9. List the number of electrons, protons and neutrons present in each atom of the following isotopes of certain elements. An example is done for you.

| | | |
|---|---|---|
| (a) | $_{6}C^{12}$ 6 protons 6 electrons 6 neutrons | $_{6}C^{14}$ 6 protons 6 electrons 8 neutrons |
| (b) | $_{17}Cl^{35}$ | $_{17}Cl^{37}$ |
| (c) | $_{1}H^{1}$ | $_{1}H^{3}$ |
| (d) | $_{92}U^{235}$ | $_{92}U^{238}$ |

10. (a) Why do we arrange the elements according to atomic number in the form of the periodic table rather than in a long horizontal row?

_____

_____

_____

(b) State the Periodic Law.

_____

_____

(c) What is the common characteristic of all the elements in group O?

_____

(d) Where do the non-metallic elements appear in the periodic table?

_____

(e) Metallic characteristics _____ from left to right in the periodic table.
(increase, decrease)

(f) List two common properties of the alkali metals.

_____

## E. Atomic Theory

All through the eighteenth century the vague idea of basic particles puzzled scientists. The actual mechanics of chemical reaction remained unsolved until John Dalton proposed an atomic hypothesis based entirely on weight relationships that had been discovered in chemical reactions. The assumptions in Dalton's theory have been revised more than once, producing a modern atomic theory, the main points of which follow.

Modern Chemical Atomic Theory

(1) A chemical element is composed of extremely small particles called atoms. Atoms of different chemical elements have different chemical properties and apparently for the most part these atoms retain their inherent difference during all ordinary chemical changes.

(Atoms of the element mercury combine with the atoms in the element oxygen to form a compound, an oxide of mercury. When this oxide of mercury is decomposed by heat, elemental mercury and elemental oxygen are obtained. This oxygen appears to be identical to the oxygen used in the preparation of the oxide and the mercury appears to be identical to the mercury used.)

(2) (a) The atoms which compose a given chemical element appear to have identical chemical properties.

(All samples of oxygen no matter where found have the chemical tendency of supporting combustion or burning.)

(b) Atoms of different chemical elements have different chemical properties.

(Dry wood will burn readily in the presence of oxygen but it will not burn solely in the presence of other non-metals such as nitrogen or sulphur.)

(3) Whenever the atoms of two or more different elements combine chemically, the atoms unite in certain simple ratios to form the smallest discrete particles of a compound.

(In the compound water, $H_2O$, hydrogen atoms and oxygen atoms are in combination in the ratio of two atoms of hydrogen to one atom of oxgen, i.e. in the atomic ratio 2:1. Atoms combine in whole numbers, never fractions.)

(4) In certain cases, the atoms of a single element unite or combine to form discrete particles. The particles are usually called molecules and each molecule of this type contains at least two identical atoms. The actual number of atoms per molecule depends upon the nature of the atoms and the mode of union.

(At ordinary temperatures hydrogen atoms unite to form diatomic molecules, $H_2$; oxygen atoms to form diatomic molecules, $O_2$; sulphur atoms to form octatomic, $S_8$, and hexatomic, $S_6$, molecules. In diamond, a kind of elemental carbon, the carbon atoms are united in a polyatomic molecule, actually a macro-molecule, whose size is that of the diamond itself. Helium molecules are monatomic, He.)

NOTE: A subscript is always used to show the number of atoms in a molecule. The subscript one, e.g. in He, is understood.

(5) The atoms of an element may combine with the atoms of another element in more than one ratio, so that more than one kind of compound may be produced.

(Carbon atoms and oxygen atoms are in combination in at least two different compounds or oxides of carbon. In carbon monoxide, CO, the carbon to oxygen ratio is 1:1; in carbon dioxide, $CO_2$, this ratio is 1:2. The ratios are in small whole numbers.)

Summary

The main points of the atomic theory are as follows:

1. The elements are made up of particles called atoms.
2. The atoms of a particular element have the same properties.
3. The atoms of different elements have different properties.
4. When atoms combine with other atoms they do so in simple ratios such as 1:1 or 2:1 or 3:2. Never do fractions of an atom combine.

Exercise 3: Atomic Theory

1. The first man to propose a widely accepted atomic theory was _____.

2. A substance containing only one atom in its molecule is called _____.

3. When the same elements combine to form two or more compounds the ratios are in the form of _____.

4. The kinetic-molecular theory cannot explain chemical change because it explains the motion of molecules which is a:
(i) physical behavior (ii) chemical behavior. (Underline the correct answer.)

5. Dalton's atomic hypothesis was based entirely on _____ _____ that had been discovered in chemical reactions.

6. The chemical elements are composed of extremely small discrete particles called _____.

7. When a compound is prepared by the union of two elements, then decomposed, the elements that have combined are _____ as the elements coming out of the reaction.

8. Atoms of different samples of the same chemical element have _____ chemical properties.

9. Atoms of different chemical elements have _____ chemical properties.

10. In the compound salt, NaCl, atoms of sodium, Na, and chlorine, Cl, have united in the ratio of _____.

11. When two atoms of chlorine combine to form a molecule $Cl_2$, the molecule is of the _____ type.

F. Gay-Lussac's Law of Volumes

At the time that Dalton proposed his atomic theory, the French scientist Gay-Lussac became interested in reactions between gases. He noticed that when water is decomposed by an electric current the volume of hydrogen liberated is exactly double that of oxygen liberated at the same time. Similarly, when hydrogen is burned in oxygen, the volume of hydrogen used and the volume of steam produced are equal and each is exactly double the volume of the oxygen used.

i.e. 2 volumes of hydrogen + 1 volume of oxygen $\longrightarrow$ 2 volumes of gaseous water

The volumes of hydrogen and oxygen and the gaseous water are in the simple ratio of

2:1:2.

Gay-Lussac also found that when hydrogen and chlorine combine to form hydrogen chloride,

1 volume of hydrogen + 1 volume of chlorine $\longrightarrow$ 2 volumes of hydrogen chloride

By means of many similar experiments he found that although the weights of the reacting gases and those produced showed no simple ratio to each other other, the volumes of these gases showed the simple ratio of small whole numbers.

These facts which could be easily verified were summed up in Gay-Lussac's Law of Combining Gas Volumes.

> At a given temperature and pressure gases combine in volumes bearing a simple ratio to one another and to the volumes of the gaseous products.

Example If 74 mL of oxygen were exploded with 28 mL of carbon monoxide to form 28 mL of carbon dioxide and 60 mL of unused oxygen, show that this experiment illustrates Gay-Lussac's Law of Volumes.

Solution: Amount of oxygen used in reaction = (74 - 60) = 14 mL
Therefore,

14 mL of oxygen react with 28 mL of carbon monoxide to yield 28 mL of carbon dioxide.

Ratio of volumes of reacting gases and product is

14:28:28 or 1:2:2,

Thus the Law is illustrated.

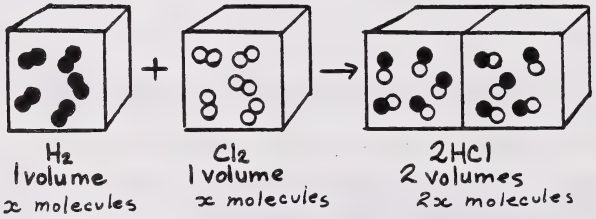
## G. Avogadro's Principle

The Italian scientist Avogadro felt that the Law of Combining Gas Volumes held a clue to the difference between atoms and molecules. He suggested that the reason why the volumes of reacting gases are in a simple ratio is that:

Equal volumes of all gases at the same temperature and pressure contain the same number of molecules.

This statement is called Avogadro's Principle.

In the reaction between hydrogen and chlorine equal volumes of the two gases form a volume of hydrogen chloride twice as great as that of either of the original gases.



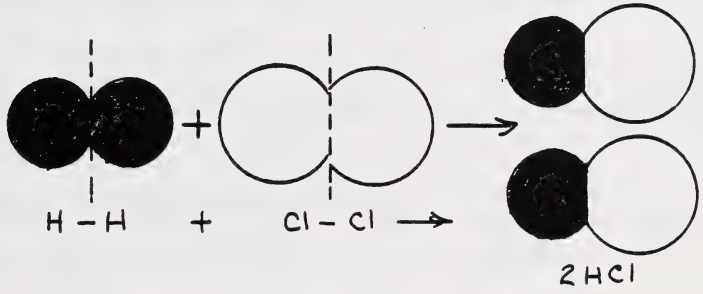

Supposing that equal volumes of hydrogen and chlorine each contain x molecules, then the first molecule of hydrogen combines with the first molecule of chlorine until the x'th molecule of the two gases combine to complete the reaction between them. Thus equal numbers of molecules have combined. But we have found experimentally that equal volumes combine.

Equal volumes of two gases must therefore contain equal numbers of molecules. This is why both the volumes and molecules are used up at the same time.

The Difference Between Molecules and Atoms

As stated previously, in the reaction between hydrogen and chlorine, the volume of the hydrogen chloride produced is twice as great as that of either the hydrogen and the chlorine. Avogadro therefore concluded that the number of molecules of hydrogen chloride produced is twice as great as the number of molecules of hydrogen and the number of molecules of chlorine used to produce it. This would mean that one molecule of chlorine and one molecule of hydrogen combine to form two molecules of hydrogen chloride.

Each of the molecules of hydrogen chloride contains one atom of hydrogen and one of chlorine. Therefore Avogadro concluded that each hydrogen particle had been split into two halves, one for each of the molecules of hydrogen chloride, similarly with each chlorine particle.




Since atoms could not be divided by chemical means, the particles that make up hydrogen gas and chlorine gas are not simple atoms but must be molecules that contain two atoms.

Summary

1. The molecule is the smallest particle that can exist physically and still retain the properties of the original pure substance.
2. The atom is the smallest particle of an element able to enter a chemical reaction.
3. Molecules that are diatomic (have two atoms to a molecule) include the gases fluorine, $F_2$; chlorine, $Cl_2$; nitrogen, $N_2$; hydrogen, $H_2$; and oxygen, $O_2$.
4. The terms atom and molecule may be used interchangeably for a monatomic molecule (molecule containing one atom) e.g. helium, He; copper, Cu; zinc, Zn; etc.)

## Avogadro's Number

Since Avogadro's time scientists have been able to determine the actual number of molecules in a given mass of an element. Twelve grams of carbon-12 (the standard) have been found to contain $6.02 \times 10^{23}$ molecules. This number is called Avogadro's number. It is also referred to as a mole of particles.

## The Avogadro Number

A large number such as $6.02 \times 10^{23}$ cannot be compared with any of the numbers we encounter in everyday life. If we counted all the people presently living on the earth it would come to only about $4 \times 10^{9}$ persons. Written out in numbers $6.02 \times 10^{23}$ is 602 followed 21 zeroes.

602 000 000 000 000 000 000 000 000

This number is much larger than any number we are accustomed to thinking about.

As a suggestion of how large $6.02 \times 10^{23}$ actually is, consider the volume of water in all the oceans and seas on the earth's surface. These contain about $1.4 \times 10^{9}$ $km^3$ (cubic km) of water. This volume is equivalent to $1.4 \times 10^{24}$ mL. If you were at the seashore and removed 2 mL of seawater, you will have removed from the total volume of ocean water a fraction of the water whose magnitude is just about the same as if you had attempted to remove one molecule of oxygen from the molecules in 32 g of oxygen.

Suppose that all the people in the world were asked to count oxygen molecules at the rate of one per second. How long would it take them to count the molecules in 32 g of oxygen?

$$\text{No. of seconds} = \frac{6.02 \times 10^{23} \text{ molecules}}{4 \times 10^{9} \text{ people}} = 1.5 \times 10^{14} \text{ s}$$

$$\text{No. of years} = \frac{1.5 \times 10^{14} \text{ s}}{(60 \times 60 \times 24 \times 365) \text{ s/yr.}}$$

$$= \frac{1.5 \times 10^{14}}{3.15 \times 10^{7}} \text{ years}$$

$$= 4.8 \times 10^{6} \text{ years}$$

$$= 4\ 800\ 000 \text{ years}$$

Exercise 4

1. (a) State the Law of Combining Volumes.

_____

_____

(b) Apply this law to (i) the decomposition of water and (ii) the union of hydrogen gas and oxygen gas to produce water.

_____

_____

_____

2. State Avogardo's Principle.

_____

_____

_____

3. The size of molecules in a gas have little to do with the volume of a gas because spaces between molecules are _____

(In any gas approximately $\frac{999}{1000}$ of it is empty space.)

4. Distinguish between an atom and a molecule, and state when the terms may be used interchangeably.

_____

_____

_____

5. (a) What is Avogadro's number?

(b) How many molecules are there in 1 g of oxygen at STP?

6. If 20 $cm^3$ of nitrogen gas react with 80 $cm^3$ of hydrogen gas to yield 40 $cm^3$ of ammonia gas and 20 $cm^3$ of hydrogen remain unreacted, show that these data illustrate Gay-Lussac's Law of Volumes.

7. When carbon monoxide and oxygen react to form carbon dioxide, 100 $cm^3$ of carbon monoxide react with 50 $cm^3$ of oxygen to produce 100 $cm^3$ of carbon dioxide. Show that this reaction illustrates Gay-Lussac's Law.

END OF LESSON C

# Chemistry Upgrading
# Lesson D

**DO NOT** send this lesson for correction

CONTENTS

A. Subscripts and Coefficients

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Table Showing Charges on Ions
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Binary Compounds

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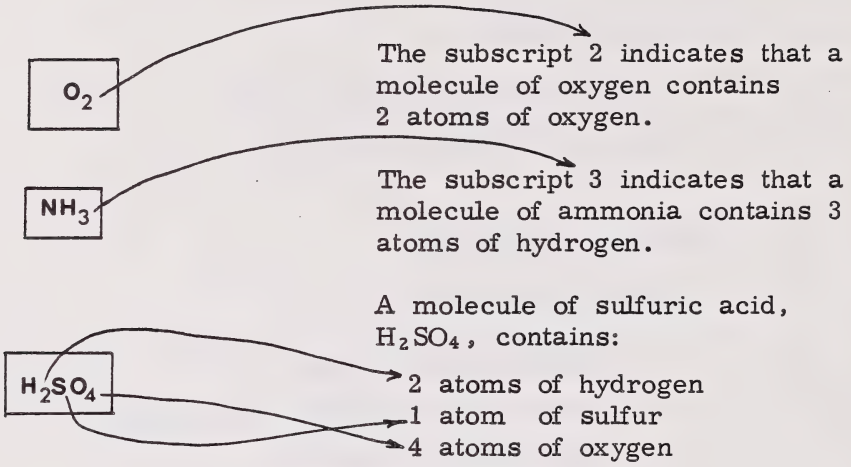
Bases
Salts
Acid Salts

D. Molecular and Atomic Masses

## A. Subscripts and Coefficients

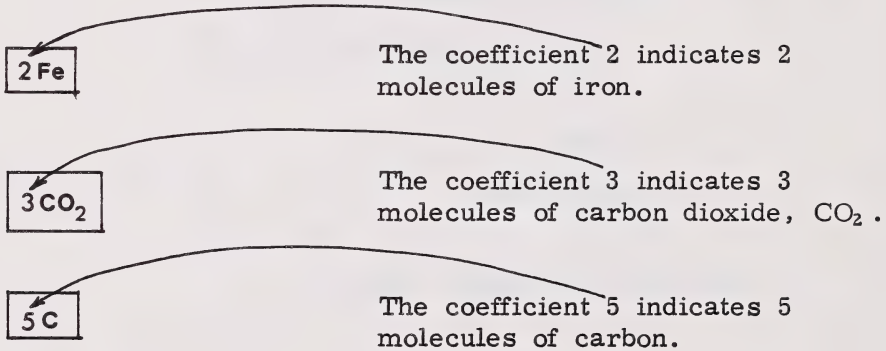
1. A SUBSCRIPT after a symbol of an element indicates the number of atoms of that element present.

e.g.

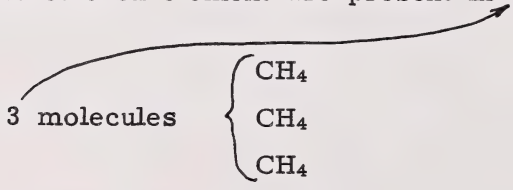

$O_2$
The subscript 2 indicates that a molecule of oxygen contains 2 atoms of oxygen.
$NH_3$
The subscript 3 indicates that a molecule of ammonia contains 3 atoms of hydrogen.
A molecule of sulfuric acid, $H_2SO_4$, contains:
$H_2SO_4$
2 atoms of hydrogen
1 atom of sulfur
4 atoms of oxygen

2. COEFFICIENTS are used to indicate the number of molecules.

e.g.


2 Fe
The coefficient 2 indicates 2 molecules of iron.
$3CO_2$
The coefficient 3 indicates 3 molecules of carbon dioxide, $CO_2$.
5 C
The coefficient 5 indicates 5 molecules of carbon.

If the coefficient indicates the number of molecules, how many atoms of each element are present in $3CH_4$?




No. of C atoms is 3

No. of H atoms is 4 + 4 + 4 = 12

## B. Writing Formulas

Chemical formulas are a shorthand method of writing the chemical name of a compound. However formulas are a more convenient way of expressing a compound than the chemical name, for, not only do they tell us the elements present in the compound but they also tell us the relative number of atoms of each element present. The name calcium chloride, for example, indicates only that the compound contains two elements, calcium and chlorine. The formula, $CaCl_2$, indicates the relative number of atoms of each element in the compound; calcium chloride contains twice as many chlorine atoms as calcium atoms. Today it is not always necessary for a chemist to analyze a compound in order to determine its formula. By knowing the combining abilities of different atoms or the charge on their ions it is possible to predict the formulas for many of the simpler compounds.

### Method of Writing Formulas

In writing the formula for any compound remember that the sum of the positive and negative charges on all the different ions present in the compound must be zero. As we saw in the previous lesson this is because the number of electrons gained by one element must be equal to the number of electrons lost by another in order for the compound to be electrically neutral. The writing of formulas is simply the balancing of the positive charge on one ion with the negative charge on another.

To write the formula for sodium chloride first write the symbols for the two types of atoms involved:

$$Na \quad Cl$$

Now, as superscripts to the right of each symbol, show the charge on each ion. (These charges can be obtained from the table on page 4 or they can be easily determined by knowing how many electrons each atom loses or gains in becoming an ion.) Since the charge on the sodium ion is +1 and the charge on the chlorine ion is -1 write:

$$Na^{+1} \quad Cl^{-1}$$

The charges on the two ions add up to zero. The formula for sodium chloride is therefore $NaCl$.

NOTE: The metallic ion, that is the ion with the positive charge, is customarily written first.

Because of the structure of their atoms, the ions of some elements have more than one possible charge. These different charges may be indicated by Roman numerals in parentheses after the name of the element.

$Cu^{+1}$ - Copper (I)
$Cu^{+2}$ - Copper (II)
$Hg^{+1}$ - Mercury (I)
$Hg^{+2}$ - Mercury (II)
$Fe^{+2}$ - Iron (II)
$Fe^{+3}$ - Iron (III)

Also see the back of the periodic table which was in Lesson B. Note the similarities along a group.

Charge on the Ions of Some Common Elements and Radicals

| Element | Metallic Ion | Element | Nonmetallic Ion | Radical | Ion |
|---|---|---|---|---|---|
| Potassium | $K^{+1}$ | Bromine | $Br^{-1}$ | Ammonium | $NH_4^{+1}$ |
| Sodium | $Na^{+1}$ | Chlorine | $Cl^{-1}$ | Chlorate | $ClO_3^{-1}$ |
| Lithium | $Li^{+1}$ | Fluorine | $F^{-1}$ | Bromate | $BrO_3^{-1}$ |
| Silver | $Ag^{+1}$ | Iodine | $I^{-1}$ | Iodate | $IO_3^{-1}$ |
| Hydrogen | $H^{+1}$ | Oxygen | $O^{-2}$ | Nitrate | $NO_3^{-1}$ |
| Barium | $Ba^{+2}$ | Sulphur | $S^{-2}$ | Nitrite | $NO_2^{-1}$ |
| Calcium | $Ca^{+2}$ | Nitrogen* | $N^{-3}$ | Acetate | $C_2H_3O_2^{-1}$ |
| Cobalt | $Co^{+2}$ | Phosphorus* | $P^{-3}$ | Hydroxide | $OH^{-1}$ |
| Magnesium | $Mg^{+2}$ | | | Cyanide | $CN^{-1}$ |
| Zinc | $Zn^{+2}$ | | | Sulphate | $SO_4^{-2}$ |
| Lead* | $Pb^{+2}$ | | | Bisulphate | $HSO_4^{-1}$ |
| Aluminum | $Al^{+3}$ | | | Sulphite | $SO_3^{-2}$ |
| Bismuth | $Bi^{+3}$ | | | Bisulphite | $HSO_3^{-1}$ |
| Copper | $Cu^{+1}$ $Cu^{+2}$ | | | Carbonate | $CO_3^{-2}$ |
| Mercury | $Hg^{+1}$ $Hg^{+2}$ | | | Bicarbonate | $HCO_3^{-1}$ |
| Iron | $Fe^{+2}$ $Fe^{+3}$ | | | Chromate | $CrO_4^{-2}$ |
| Chromium | $Cr^{+2}$ $Cr^{+3}$ | | | Phosphate | $PO_4^{-3}$ |
| Tin | $Sn^{+2}$ $Sn^{+4}$ | | | Aluminate | $AlO_3^{-3}$ |

*Other charges on ions are possible.

## Rules for Writing Formulas

The total negative charge must be equal to the total positive charge in order that the molecule be electrically neutral.

(a) When ions with charges of the same numerical value combine, one atom of each element will produce a balance of charge.

| Examples: | Metallic Ions | Nonmetallic Ions | Formula |
|---|---|---|---|
| | $Na^{+1}$ | $C\ell^{-}$ | $NaC\ell$ |
| | $Ca^{+2}$ | $O^{-2}$ | $CaO$ |
| | $A\ell^{+3}$ | $N^{-3}$ | $A\ell N$ |

(b) When the ions have different charges there must be more atoms of the element whose ion has the lower charge.
e.g. the formula for calcium bromide
Write the symbols for the two ions placing the charge as a superscript to the right of each symbol.

$$Ca^{+2}\ Br^{-1}$$

The charges do <u>not</u> add up to zero therefore CaBr is <u>not</u> the correct formula. Because each calcium atom loses 2 electrons and each bromine atom accepts only 1, two atoms of bromine are needed to accept the electrons from a single calcium atom. The combination of 1 atom of calcium with 2 atoms of bromine is written by placing the subscript 2 to the right of the bromine:

$$CaBr_2$$

To check the balance of charge:

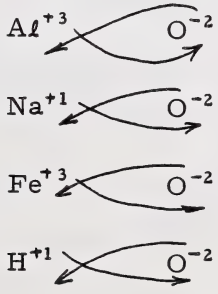
$$Ca^{+2}Br_2{}^{-1}$$

positive charge = $(+2) \times 1 = +2$
negative charge = $(-1) \times 2 = -2$
Charge on compound = $+2 - 2 = 0$

## Simplified Rule

To obtain the formula crisscross the charges on the ions (ignoring the signs).

Examples:




Formula
$A\ell_2O_3$
$Na_2O$
$Fe_2O_3$
$H_2O$

## Radicals

In some compounds atoms are found combined in the form of radicals. These are ions consisting of more than one type of atom. They usually remain unchanged during a chemical reaction because they act as a unit having a fixed charge. See the table on page 4 for the formulas and charges of the more common radicals.

When writing the formula for a compound containing a radical consider the radical as a single unit with a predetermined formula and charge. Never separate the radical into its individual elements.

To write the formula for magnesium sulphate note that the charge on the magnesium ion is +2 and the charge on the sulphate ion is -2. Therefore written together they become:

$$Mg^{+2}SO_4^{-2}$$

Since the +2 charge of the magnesium ion and the -2 charge of the sulphate ion add up to zero, the formula for magnesium sulphate is $MgSO_4$.

To write the formula for aluminum sulphate show the ionic charges

$$Al^{+3}SO_4^{-2}$$

By the simplified crisscross method of placing the number 2 as a subscript to the right of the aluminum ion and the subscript 3 to the right of the sulphate ion, the formula becomes:

$$Al_2^{+3}(SO_4)_3^{-2}$$

The total ionic charge adds up to zero.

$$[2 \times (+3)] + [3 \times (-2)] = 0$$

Therefore the correct formula is $Al_2(SO_4)_3$.

Note that parentheses are placed around the $SO_4$ to show that the subscript 3 refers to the whole sulphate radical. The parentheses also prevent confusion between the subscript and any number in the radical. Example:

| | | Formula |
|---|---|---|
| zinc chlorate | $Zn^{+2}ClO_3^{-1}$ | $Zn(ClO_3)_2$ |

## Exercise 1 - Writing Formulas

1. With the use of the enclosed ionic table on page 4, write the formula for the compounds containing (a) pairs of elements and (b) elements and radicals. Remember that the net charge on the compound must be zero since the compound is electrically neutral. Several examples are done for you.

(a)

| Pairs of Elements | Ions and Charge | Formula |
|---|---|---|
| mercury (I), oxygen | $Hg^{+1}O^{-2}$ | $Hg_2O$ |
| silver, sulphur | | |
| potassium, iodine | | |
| copper (I), sulphur | | |
| aluminum, sulphur | | |
| calcium, chlorine | | |
| iron (III), oxygen | | |
| zinc, bromine | | |
| barium, chlorine | | |
| lithium, fluorine | | |
| sodium, oxygen | | |
| calcium, iodine | | |
| mercury (II), oxygen | $Hg^{+2}O^{-2}$ | HgO not $Hg_2O_2$ |
| iron (II), oxygen | | |

(b)

| Element and Radical | Ions and Charge | Formula |
|---|---|---|
| sodium carbonate | $Na^{+1}CO_3^{-2}$ | $Na_2CO_3$ |
| barium nitrate | | |
| aluminum hydroxide | | |
| mercury (I) nitrate | | |
| iron (II) sulphate | | |
| iron (III) sulphate | | |
| lithium hydroxide | | |
| ammonium carbonate | | |
| calcium acetate | | |
| aluminum bisulphite | | |
| sodium cyanide | | |
| chromium (II) phosphate | | |

## C. Naming of Compounds

The system of naming compounds is designed to supply information concerning the elements contained in the compound.

### Binary Compounds

Binary compounds consist of two elements only. The metallic or positive part is written first; then the nonmetallic or the negative part. The name of a binary compound ends in "ide".

$NaCl$ - sodium chlor<u>ide</u>
$CaO$ - calcium ox<u>ide</u>
$H_2S$ - hydrogen sulph<u>ide</u>

### Oxides and Peroxides

An oxide is a binary compound containing the element oxygen.

$Na_2O$ - sodium <u>oxide</u>
$BaO$ - barium <u>oxide</u>
$Al_2O_3$ - aluminum <u>oxide</u>

Note that the above examples conform to the charge rules. They are called the normal metallic oxides. There are some oxides which do not conform to the charge rules. They possess one more atom of oxygen per molecule than the normal oxide. Such oxides are designated by the prefix "per".

$Na_2O_2$ - sodium peroxide
$BaO_2$ - barium peroxide
$H_2O_2$ - hydrogen peroxide

## Metals With Two Possible Charges

When a metallic ion can have two different charges, write a Roman numeral indicating the charge, after the name of the metal.

| | | |
|---|---|---|
| e.g. | FeO | iron II oxide |
| | $Fe_2O_3$ | iron III oxide |
| | $Cu_2O$ | copper I oxide |
| | CuO | copper II oxide |
| | HgBr | mercury I bromide |
| | $HgBr_2$ | mercury II bromide |

## Nonmetallic Compounds

Sometimes nonmetals act as metals in the formation of compounds; that is, they combine with other nonmetals. In the naming of such compounds, the suffixes "ous" and "ic" are not preferred. Prefixes such as "mon" or "mono", "di", "tri", "tetra", "pent" or "penta" etc. are used before the name of the elements to indicate the number of atoms of the element.

| | |
|---|---|
| CO | carbon monoxide |
| $CO_2$ | carbon dioxide |
| $SO_2$ | sulphur dioxide |
| $SO_3$ | sulphur trioxide |
| $CC\ell_4$ | carbon tetrachloride |
| $P_2O_3$ | diphosphorus trioxide |
| $P_2O_5$ | diphosphorus pentoxide |

Note: mon or mono - one
di - two
tri - three
tetra - four
pent or penta - five

## Binary Acids

Binary acids are made up of two elements, one of which is hydrogen. If the compound is in the gaseous state, its name ends in "ide". The acids are produced by dissolving the corresponding gases in water. The name of a binary acid begins with the prefix "hydro" and ends with the suffix "ic".

| Gaseous Compound | Acid |
|---|---|
| HF - hydrogen fluoride | HF - hydrofluoric acid |
| HCl - hydrogen chloride | HCl - hydrochloric acid |
| HBr - hydrogen bromide | HBr - hydrobromic acid |
| HI - hydrogen iodide | HI - hydriodic acid |
| $H_2S$ - hydrogen sulphide | $H_2S$ - hydrosulphuric acid |

## Oxy-Acids or Ternary Acids

Oxy-acids consist of three elements, two of which are hydrogen and oxygen. The standard oxy-acid ends with "ic" e.g. $H_2SO_4$ - sulphuric acid - but often the number of oxygen atoms in a molecule of acid varies. As the following oxy-acids are very common in all chemical work you are advised to memorize their names and formulas.

| Oxy-Acid | Formula |
|---|---|
| nitric | $HNO_3$ |
| chloric | $HClO_3$ |
| sulphuric | $H_2SO_4$ |
| carbonic | $H_2CO_3$ |
| phosphoric | $H_3PO_4$ |
| bromic | $HBrO_3$ |
| iodic | $HIO_3$ |

If a molecule of an oxy-acid contains one more atom of oxygen than the standard "ic" acid, this is designated by the prefix "per" and the suffix "ic".

If the molecule contains one atom of oxygen less than the "ic" acid, it is designated by the suffix "ous".

If the molecule contains two atoms of oxygen less than the "ic" acid, it is designated by the prefix "hypo" and the suffix "ous".

$HClO_4$ - perchloric acid

$HClO_3$ - chloric acid

$HClO_2$ - chlorous acid

$HClO$ - hypochlorous acid

## Bases

The molecule of a base consists of a metal or metallic radical and the hydroxyl radical, $OH^-$. The hydroxyl radical is nonmetallic in nature (charge of -1) and therefore appears at the end.

Bases are named hydroxides.

| | | |
|---|---|---|
| LiOH | - | lithium hydroxide |
| $Ca(OH)_2$ | - | calcium hydroxide |
| $NH_4OH$ | - | ammonium hydroxide |

## Salts

There are several ways in which salts may be formed. A salt is produced when the hydrogen of an acid is displaced by a metal or a metallic radical. Binary salts are formed from binary acids. Like all binary compounds, the names of the binary salts end in "ide".

| Binary Acid | | | Binary Salt | | |
|---|---|---|---|---|---|
| | | H replaced by Na | | | |
| $HCl$ | - | hydrochloric acid | $NaCl$ | - | sodium chloride |
| | | H replaced by K | | | |
| HBr | - | hydrobromic acid | KBr | - | potassium bromide |
| HI | - | hydriodic acid | $CaI_2$ | - | calcium iodide |
| $H_2S$ | - | hydrosulphuric acid | CuS | - | cupric sulphide |

Oxy-acids form oxy-salts. The name of the salt has a definite relationship to the name of its corresponding acid.

A "per .... ic" acid forms a "per .... ate" salt.

An "ic" acid forms an "ate" salt.

An "ous" acid forms an "ite" salt.

An "hypo .... ous" acid forms an "hypo .... ite" salt.

| Acid | | | Salt | | |
|---|---|---|---|---|---|
| $HBrO_4$ | - | perbromic acid | $KBrO_4$ | - | potassium perbromate |
| $HBrO_3$ | - | bromic acid | $KBrO_3$ | - | potassium bromate |
| $HBrO_2$ | - | bromous acid | $KBrO_2$ | - | potassium bromite |
| HBrO | - | hypobromous acid | KBrO | - | potassium hypobromite |

Acid Salts

Salts are formed when all the replaceable hydrogen of an acid is displaced by a metal or metallic radical.

$H_2SO_4$ - sulphuric acid

$Na_2SO_4$ - sodium sulphate

Acid salts are formed when only part of the replaceable hydrogen of an acid is displaced by a metal or a metallic radical. Acid salts are named by using the prefix "bi" before the nonmetallic radical.

$NaHSO_4$ - sodium bisulphate

$NaHSO_3$ - sodium bisulphite

$KHCO_3$ - potassium bicarbonate

The prefix "bi" is used only in the naming of acid salts which contain nonmetallic radicals with a charge of (-2). (i.e. $SO_4^{-2}$, $CO_3^{-2}$ etc.) One of the two replaceable hydrogens is displaced by a metal or metallic radical.

In the case where the salt is the product of an acid containing a nonmetallic radical with a charge of (-3) and three atoms of hydrogen e.g. $H_3PO_4$, the prefix "bi" is not used since either one or two hydrogen atoms may be displaced to form an acid salt. i.e. Na may displace either one or two of the displaceable hydrogens in $H_3PO_4$ to form either $Na_2HPO_4$ or $NaH_2PO_4$. The prefixes "mono" or "di" are used in this case.

$Na_2HPO_4$ - disodium phosphate or
sodium monohydrogen phosphate

$NaH_2PO_4$ - monosodium phosphate or
sodium dihydrogen phosphate

Exercise 2

1. Write the formulas for the following compounds.

   (a) sodium peroxide _____

   (b) hydrosulphuric acid _____

   (c) sulphuric acid _____

   (d) sulphurous acid _____

   (e) perchloric acid _____

   (f) carbon monoxide _____

   (g) phosphorus trioxide _____

2. Name the following compounds.

   (a) HI _____

   (b) NaI _____

   (c) $KIO_3$ _____

   (d) $HIO_3$ _____

   (e) $H_2O_2$ _____

   (f) HgO _____

   (g) $MgSO_4$ _____

   (h) MgS _____

   (i) $A\ell C\ell_3$ _____

   (j) KCN _____

(k) $KClO_4$ _____

(l) $NaHSO_3$ _____

(m) $Fe(OH)_3$ _____

(n) $CoCl_2$ _____

(o) $ZnCO_3$ _____

Exercise 3

1. Match the following compounds with the description that best fits them. An example is done for you.

| | | |
|---|---|---|
| (a) $H_2S$ | d | (a) binary salt ending in "ide" |
| (b) $H_2SO_4$ | _____ | (b) iron compound in which iron ion has a charge of +2 |
| (c) $NaH_2PO_4$ | _____ | (c) iron compound in which iron has a charge of +3 |
| (d) $KCl$ | _____ | (d) binary acid "hydro .... ic" |
| (e) $Fe_2(SO_4)_3$ | _____ | (e) acid salt |
| (f) $H_2O_2$ | _____ | (f) ternary "ite" salt |
| (g) $P_2O_5$ | _____ | (g) compound formed by two nonmetallic elements |
| (h) $Ca(ClO_4)_2$ | _____ | (h) oxy-acid ending in "ic" |
| | | (i) a "per .... ate" oxy-salt |
| | | (j) a "hypo .... ite" oxy-salt |
| | | (k) a "per" oxide -- one more oxygen than normal |

2. (a) Name and write the formulas for the oxides of the following metals. An example is done for you.

copper (I) - $Cu_2O$ - copper I oxide

copper (II) - $CuO$ - copper II oxide

mercury (I) - _____

mercury (II) - _____

iron (II) - _____

iron (III) - _____

tin (II) - _____

tin (IV) - _____

(b) Name and write the formulas for the two oxides of the following nonmetals.

carbon _____

_____

phosphorus _____

_____

3. The rare element selenium forms a ternary acid which greatly resembles sulphuric acid. The acid is named selenic acid and its formula is $H_2SeO_4$. Write the formulas and names of the normal salts of this acid with the following metals.

The charge on the selenate ion is 2-.

silver $Ag_2SeO_4$ silver selenate

calcium _____

iron (II) _____

iron (III) _____

tin (II) _____

tin (IV) $Sn(SeO_4)_2$ tin (IV) selenate

4. Name the chemical substances represented by the following chemical formulas.

(a) $H_3PO_2$ _____

(b) $Pb(OH)_2$ _____

(c) $CO$ _____

(d) $CO_2$ _____

(e) $NH_3$ _____

4. (f) $NH_4Br$ _____

(g) $PBr_5$ _____

(h) $P_2O_3$ _____

(i) $(NH_4)_2SO_3$ _____

(j) $Na_2HPO_4$ _____

A number of gases, when they appear by themselves, are always in the form of diatomic gases. There are only 8 diatomic molecules. They are: hydrogen ($H_2$), nitrogen ($N_2$), oxygen ($O_2$), and the five halogens of group VIIA: $F_2$, $Cl_2$, $Br_2$, $I_2$, $At_2$. Whenever one of these eight elements combines with anything else, always write these elements in diatomic form when they appear alone, never as a single atom. However, when these elements appear in a compound they are never in a compound as a diatomic gas. The following illustration should make this clear. When magnesium (Mg) reacts with oxygen, ($O_2$) magnesium oxide (MgO) is produced. To write this with symbols, it would be

$$Mg + O_2 \longrightarrow MgO$$

Do <u>not</u> write:

$Mg + O \longrightarrow MgO$, ×

nor $Mg + O_2 \longrightarrow MgO_2$ ×

## D. Molecular and Atomic Masses

### Method for Finding Molecular Masses of Gases

Weighing a single molecule or atom on a balance is impossible. However, it is a fairly simple matter, to weigh equal volumes of different gases at the same temperature and pressure. In effect, this is the same as weighing equal numbers of molecules of the various gases, because the ratio of the weights of the equal volumes is equal to the ratio of the weights of the molecules. Avogadro's law spares us the need of having to weigh single molecules on a balance.

If, for example, a certain gas A contains n molecules, then an equal volume of gas B would also contain n molecules at the same temperature and pressure (Avogadro's Principle). Suppose we find that gas A weighs twice as much as gas B. It can then be concluded that one molecule of A weighs twice as much as one molecule of B.

As a more realistic example, suppose a certain volume of oxygen at 0°C and 760 mm of pressure has a mass of 32 g. The same volume of hydrogen under the same conditions has a mass of 2 g. If the number of molecules of oxygen is y, the number of hydrogen molecules would also be y. Since the mass of y molecules of oxygen is 16 times as much as the mass of y molecules of hydrogen the mass of one molecule of oxygen is 16 times as much as the mass of one molecule of hydrogen.

## Standard for Atomic and Molecular Masses

The above idea has been used by chemists to find the relative masses of molecules of gases, and of the atoms which form such molecules. To apply the method effectively it was necessary to choose an atom to which all other atoms could be compared. The atom of the most common isotope of carbon was chosen for this purpose. It has been assigned a value of exactly 12.000 atomic mass units. Then the atomic mass of oxygen is 16 atomic mass units. Since oxygen is diatomic (2 atoms to the molecule) the mass of a molecule of oxygen is 32 a.m.u.

## The Atomic Mass Unit

The standard of atomic mass is the isotope of carbon of mass 12. The most abundant isotope is assigned a mass of 12 atomic mass units (a.m.u.). Thus, one a.m.u. is 1/12 the mass of the "standard" isotope of carbon.

With the atomic mass unit it becomes possible to assign units to the atomic and molecular mass. The statement that "the atomic mass of carbon is 12" is equivalent to "the mass of carbon's atom is 12 a.m.u." Similarly the atomic masses of oxygen and hydrogen, in round numbers would be 16 a.m.u. and 1 a.m.u. respectively.

## The Gram and the Mole of Atoms

Because the atomic masses are relative masses, they may be expressed as absolute masses using units such as grams. The molar mass of an element is the atomic mass expressed in grams. This is also called the mass of a mole of atoms. The atomic mass of carbon is 12 a.m.u.; therefore its molar mass is 12 g or a mole of atoms has a mass of 12 g. The atomic mass of oxygen is 16 a.m.u.; therefore its molar mass is 16 g or a mole of atoms has a mass of 16 g.

As stated previously, we have chosen carbon as our standard for atomic mass with an assigned mass of 12.000 a.m.u. If we take a certain number of carbon atoms and an equal number of oxygen atoms the mass of the oxygen atoms is 16/12 times as great as the mass of the carbon atoms. In other words, any sample of oxygen that has mass 16/12 as much as a sample of carbon must contain the same number of oxygen atoms as the sample of carbon atoms. Therefore 16 g of oxygen would contain the same number of atoms as 12 g of carbon. Similarly 1 g of hydrogen would contain the same number of atoms as 12 g of carbon. This number has been determined experimentally and found to be Avogadro's number. Therefore there are $6.02 \times 10^{23}$ atoms/mol.

Table

| Element | Atomic Mass in a.m.u. | Molar Mass | Grams in 1 mol | No. of atoms in 1 mol of atoms |
|---|---|---|---|---|
| Hydrogen | 1 a.m.u. | 1 g | 1 g | $6.02 \times 10^{23}$ |
| Carbon | 12 a.m.u. | 12 g | 12 g | $6.02 \times 10^{23}$ |
| Sulphur | 32 a.m.u. | 32 g | 32 g | $6.02 \times 10^{23}$ |
| Zinc | 65 a.m.u. | 65 g | 65 g | $6.02 \times 10^{23}$ |

The number of moles in a given mass of any element can be formulated as follows:

$$\text{No. of moles} = \frac{\text{mass of element in grams}}{\text{number of grams/mole of atoms}} = \frac{\text{mass}}{\text{molar mass}}$$

To find the number of moles of atoms in a given mass of an element divide the given number of grams of the element by the molar mass of the element.

Note: mol is the abbreviation for mole.
a.m.u. is the abbreviation for atomic mass unit.

Example 1

How many moles of atoms of carbon are there in 64.00 g of carbon?

Solution:

There are 12.01 g of carbon in 1 mol of atoms of the element and 64.00 g of carbon in $\frac{64.00 \text{ g}}{12.01 \text{ g/mol}} = 5.3$ mol

NOTICE THE UNITS

**Example 2**

How many grams of carbon are there in 0.375 mol of atoms of carbon?

From the periodic table 1 mol of atoms of carbon has mass 12.01 g

$\therefore$ 0.375 mol of atoms of carbon has mass:

$$= 0.375 \text{ mol} \times \frac{12.01 \text{ g}}{\text{mol}} = \underline{4.50 \text{ g}}$$

**Relationship between Atomic Mass Units and Grams**

$6.02 \times 10^{23}$ atoms of carbon has a mass of 12.01 g

1 atom of carbon has mass $\frac{12.01}{6.02 \times 10^{23}}$ g

or $2 \times 10^{-23}$ g

The atomic mass unit is $\frac{1}{12}$ the mass of the carbon atom;

therefore 1 a.m.u. $= \frac{12.01 \text{ g}}{6.02 \times 10^{23}} \times \frac{1}{12} = \underline{1.66 \times 10^{-24} \text{ g}}$

**Exercise 4**

1. The following elements are in the atmosphere: oxygen, argon, krypton, and nitrogen. Give symbols for 1 mol of molecules for each element. (Hint: some elements are diatomic; others are monatomic.)

_____ _____ _____ _____

2. Using appropriate symbols, write the reaction when calcium combines with oxygen to produce calcium oxide.

_____

3. What element is used as the standard for atomic masses?

_____

4. What is an atomic mass unit? The atomic mass of a sulphur atom is 32 a.m.u. What does this mean?

_____

_____

5. (a) What is a mole of atoms? _____

_____

_____

(b) Find the number of moles of atoms of nitrogen in 56 g of the element.

6. Find the number of grams in 0.30 mol of atoms of helium.

7. Find the mass in grams of the following number of moles of atoms of each element. An example has been done for you showing one method of solution.

(a) 0.200 mol of atoms of sodium

Mass of 1 mol of atoms of sodium is 22.99 g

$\therefore$ mass of 0.200 mol

= 0.200 mol x 22.99 g/mol

= 4.60 g

7. (b) 1.6 mol of atoms of zinc.

(c) 4.5 mol of atoms of chlorine.

8. Calculate the mass of one mole of each of the following:

(a) $Ca(NO_3)_2$

(b) $Mn_2O_5$

8. (c) $Fe(C_2H_3O_2)_3$

9. The mass of a proton or neutron is $1.66 \times 10^{-24}$ g. Since the atomic mass of hydrogen is 1.01, it means that each hydrogen atom contains only one proton but no neutrons. (The mass of the electron is negligible.) How many hydrogen atoms would be needed to make 1.0 g of hydrogen?

HINT: $\frac{X \text{ atoms}}{1 \text{ atom}} = \frac{1.0 \text{ g}}{1.66 \times 10^{-24} \text{ g}}$ Find X

10. One mole of any substance has $6.02 \times 10^{23}$ molecules.
(a) How many molecules are there in 2 mol of any substance?

(b) How many molecules are there in 0.5 mol of any substance?

10. (c) One mole of $MgSO_4$ has $6.02 \times 10^{23}$ molecules. Each molecule of $MgSO_4$ has 6 atoms. Is this right? _____ How many <u>atoms</u> are there in 1 mol of $MgSO_4$?

(d) How many molecules are there in 0.01 mol of $NaOOCC_{17}H_{35}$? (This is the formula for a soap molecule.)

END OF LESSON D

# Chemistry Upgrading
# Lesson E

**DO NOT** send this lesson for correction

# CHEMICAL EQUATIONS

A. Writing Chemical Equations
   1. Qualitative
   2. Quantitative

B. Balancing of Chemical Equations
   1. Using Moles of Atoms
   2. Using Atoms
   3. Conservation of Mass

C. Using Equations in Problem Solving

D. Types of Chemical Reactions
   1. Synthesis Reaction
   2. Decomposition Reaction
   3. Rearrangement or Replacement Reaction
   4. Ion-ion Reaction

## A. Writing Chemical Equations

Chemical equations may be either (1) qualitative or (2) quantitative depending on the information that is available.

1. Rules to follow in writing a qualitative chemical equation. It is very important to know what is reacting (reactants) and the substances produced in the reaction (products). Also one must know the correct formulas for the reactant(s) and for the product(s) before writing the equation. As a rule, the formula for the reactants appears on the left side and the formula for the products on the right side of the arrow.

At this point it is not necessary to know how many moles or how many grams of reactants take part in the reaction, nor is it important to know how much of the product is formed in the reaction. These quantities will be dealt with later in the lesson. Also we will not worry about balancing equations now, therefore coefficients will be ignored in the reactions that follow.

You may be quite aware of the many chemical reactions taking place about you, but you really had no opportunity to express them in equation form. To think of a few examples, consider:

1. (a) Silver cutlery tarnishes in your home. Why does this occur?

Silver combines with oxygen in the atmosphere to produce a dark coating of silver oxide. The equation for this reaction is:

$$Ag + O_2(g) \longrightarrow Ag_2O$$

(b) Charcoal (carbon) burns in your barbeque to produce carbon dioxide. Equation:

$$C + O_2(g) \longrightarrow CO_2(g)$$

(c) A candle burns on your birthday cake to produce carbon dioxide and water. Equation:

$$C_{25}H_{52} \text{ (paraffin)} + O_2(g) \longrightarrow CO_2(g) + H_2O$$

(d) Baking soda neutralizes an acid stomach. Equation:

$$NaHCO_3 + HCl \text{ (acid in stomach)} \longrightarrow H_2O + CO_2(g) + NaCl$$

The above reactions are examples of qualitative chemical equations.
(Diatomic elements: $H_2$, $N_2$, $O_2$, $F_2$, $Cl_2$, $Br_2$, $I_2$, $At_2$)

Exercie 1

Using the rules given in this lesson for writing equations and using examples above, please write equations for the following reactions:

1. A copper cooking utensil tarnishes in air to produce a dark coating of copper II oxide.

_____

2. Gasoline ($C_8H_{18}$) burns completely in your automobile engine to produce carbon dioxide and water.

_____

3. Iron rusts to produce iron II oxide.

_____

4. Sodium metal reacts with water to produce hydrogen and sodium hydroxide.

_____

5. Potassium combines with bromine to yield potassium bromide.

_____

6. In flash bulbs, magnesium burns in oxygen to produce a white solid magnesium oxide.

---

7. Nitrogen combines with hydrogen (under special conditions) to produce ammonia ($NH_3$).

---

8. Chlorine reacts with sodium to produce common table salt.

---

Qualitative chemical equations may also show energy as part of the reaction. For example, water may be decomposed into hydrogen and oxygen using electrical energy.

$$H_2O(\ell) + \text{energy} \longrightarrow H_2(g) + O_2(g)$$

The above represents an endothermic reaction because energy is absorbed during reaction.

A reaction may show energy is released as in burning charcoal. We obtain energy in terms of heat.

$$C + O_2 \longrightarrow CO_2 + \text{energy}$$

The above reaction illustrates an exothermic reaction. Again we shall not discuss the quantity of energy absorbed or released since we are dealing with only qualitative aspects of a chemical reaction.

Exercise 2

Write an equation to show decomposition of potassium chlorate ($KClO_3$) by heating it strongly to produce oxygen and potassium chloride.

---

2. Rules to follow in writing Quantitative chemical equations. Here again you will be expected to know the correct molecular formula for the reactants and products but in addition you will have to show that (1) atoms are conserved and that (2) mass is conserved.

To meet these requirements you will have to balance your equation. Balancing of equations involves a simple bookkeeping device in which one keeps track of atoms taking part in a chemical reaction. The number of atoms entering a reaction must equal the number of atoms forming the new products. One may count these atoms by using the convenient mole concept. In other words, we will count moles of atoms in a given chemical reaction.

If you are given the number of moles of reactants and the number of moles of products then it would be quite simple to count the number of moles of atoms reacting and the number of moles of atoms (of each element) produced by the given reaction. For example: you may be given this:

> Two moles of red mercury II oxide are heated to produce one mole of oxygen and two moles of mercury. The above can be written as a chemical equation: $2\ HgO + \text{energy} \longrightarrow O_2 + 2\ Hg$

Is it possible to determine the number of moles of atoms reacting or the number of moles of atoms produced by observing the above equation?

Details regarding this procedure will be discussed in the next topic under Balancing of Equations.

Exercise 3

Given that one mole of propane ($C_3H_8$) reacts with five moles of oxygen to yield three moles of carbon dioxide and four moles of water, write the chemical equation for this reaction showing the number of moles of each compound involved.

---

B. Balancing of Chemical Equations

1. For the balancing of equations using moles of atoms, study the examples below.

Magnesium combines with oxygen in the air to produce a white solid, magnesium oxide. To write the equation, you must know the formula for magnesium, oxygen and magnesium oxide.
Therefore you will have:

$$Mg + O_2 \longrightarrow MgO$$

In order to conserve moles of atoms we must find numerical coefficients to place before each formula to obtain the same number of moles of atoms of each element on the left and on the right of the arrow. That is, moles of magnesium atoms on the left side must equal moles of

magnesium atoms on the right side of the arrow. The same must be true for all other elements. The process of finding these coefficients is called balancing the chemical equation.

Suppose one mole of oxygen gas is consumed in the reaction.

$Mg + 1O_2 \longrightarrow MgO$

But one mole of oxygen gas contains two moles of oxygen atoms. On your right side, one mole of magnesium oxide contains only one mole of oxygen atoms. To balance oxygen, place a 2 in front of MgO which will now give two moles of oxygen atoms.

$Mg + O_2 \longrightarrow 2MgO$

(With no coefficient, a 1 is understood to be in front.)

But now magnesium is not balanced. Place a 2 in front of Mg. Are all atoms conserved now? _____

$2Mg + O_2 \longrightarrow 2MgO$

You can use fractions to balance equations. The same equation is correctly balanced as shown below:

$$Mg + 0.5O_2 \longrightarrow MgO$$

Are moles of atoms balanced? _____

SUMMARY

$2\ Mg + O_2 \longrightarrow 2\ MgO$

$Mg + 0.5\ O_2 \longrightarrow MgO$

Either way of writing the balanced equation for the burning of magnesium is correct.

If you wish to eliminate fractions you may multiply the second equation by 2.

$2 \times (1\ Mg + 0.5\ O_2 \longrightarrow 1\ MgO) =$

$2\ Mg + 1O_2 \longrightarrow 2\ MgO$ to obtain the first equation.

Please Note: In writing chemical equations we often drop the coefficient 1, but it is never wrong to write it in the equation.

2. For the balancing of equations using atoms, study the examples below.

Given that hydrogen burns in air to yield water, your equation for the above reaction is as follows:

$$H_2 + O_2 \longrightarrow H_2O$$

| | | | | |
|---|---|---|---|---|
| 1 molecule of hydrogen gas or 2 atoms of hydrogen | + | 1 molecule of oxygen gas or 2 atoms of oxygen | ⟶ | 1 molecule of water which contains 2 atoms of hydrogen plus one atom of oxygen |

To correct the deficiency of oxygen atoms on the right side of your equation you may place the coefficient 2 before the formula $H_2O$.

Now you have:

$$H_2 + O_2 \longrightarrow \underline{2}\ H_2O$$

In 2 molecules of water there are 2 atoms of oxygen and 4 atoms of hydrogen. To correct the deficiency of hydrogen atoms on the left side you use the coefficient 2 before the formula $H_2$.

Now you have:

$$\underline{2}\ H_2 + O_2 \longrightarrow 2\ H_2O$$

Make a final check to see whether atoms are conserved.

| Left side | element | Right side |
|---|---|---|
| $2 \times 2 = 4$ | H | $2 \times 2 = 4$ |
| $1 \times 2 = 2$ | O | $2 \times 1 = 2$ |
| Total $\boxed{6}$ | | Total $\boxed{6}$ |

Atoms are conserved!

Exercise 4

Balance the following equation:

$$O_2 + Cl_2 \longrightarrow Cl_2O$$

In balancing atoms of oxygen, the equation becomes

$$O_2 + Cl_2 \longrightarrow \_\_\_\ Cl_2O$$

| Left side | element | Right side |
|---|---|---|
| $1 \times 2 = 2$ | O | $2 \times 1 = 2$ |

The equation requires correction for chlorine atoms.

The balanced equation becomes:

$$O_2 + \_\_Cl_2 \longrightarrow \_\_Cl_2O$$

Are all atoms conserved? _____ If they are, show proof below.

| Left side | element | Right side |
|---|---|---|
| | O | |
| | Cl | |
| Total ___ | | Total ___ |

Exercise 5

Acetylene burns in air to form carbon dioxide and water, represented by the following unbalanced equation:

$$C_2H_2 + O_2 \longrightarrow CO_2 + H_2O$$

First balance carbon atoms by placing the correct coefficient before the formula $CO_2$.

$$C_2H_2 + O_2 \longrightarrow \_\_CO_2 + H_2O$$

Next balance oxygen atoms by placing the correct coefficient before $O_2$ (fractions may be used).

$$C_2H_2 + \_\_O_2 \longrightarrow \_\_CO_2 + H_2O$$

To clear the fraction multiply the equation by _____. The equation becomes: _____.

Are all atoms conserved? _____

Exercise 6

Balance the following chemical equations:

(a) ___$C_2H_2$ (acetylene) + ___$O_2$ ⟶ ___$CO_2$ + ___$H_2O$

(b) ___$HgO$ ⟶ ___$Hg$ + ___$O_2$

(c) ___$Cu$ + ___$S_8$ ⟶ ___$CuS$

(d) ___$Cl_2$ + ___$NaI$ ⟶ ___$I_2$ + ___$NaCl$

(e) ___$NaOH$ + ___$H_2SO_4$ ⟶ ___$H_2O$ + ___$Na_2SO_4$

(f) ___$Li$ + ___$AlCl_3$ ⟶ ___$Al$ + ___$LiCl$

(g) ___$FeCl_3$ + ___$KOH$ ⟶ ___$Fe(OH)_3$ + ___$KCl$

(h) ___$C_3H_8$ + ___$O_2$ ⟶ ___$H_2O$ + ___$CO_2$

(i) ___$NaCl$ ⟶ ___$Na$ + ___$Cl_2$

(j) ___$K$ + ___$H_2O$ ⟶ ___$H_2$ + ___$KOH$

## 3. Conservation of Mass

A balanced equation also shows that mass is conserved. During a chemical reaction mass is neither created nor destroyed but is changed from one form to another. To illustrate the mass is conserved in a chemical reaction, consider the balanced equation for the burning of magnesium.

$$2\ Mg + O_2 \longrightarrow 2\ MgO$$

| | | | | |
|---|---|---|---|---|
| 2 mol of magnesium, Mg | + | 1 mol of oxygen, $O_2$ | ⟶ | 2 mol of magnesium oxide, MgO |

In grams, one mole of Mg = 24.3 g
∴ 2 mol of Mg = 48.6 g

one mole of $O_2$ = 32.0 g ⟶ one mole of MgO = 40.3 g, so two moles = 80.6 g

Mass of Reactants = Mass of Products

48.6 g + 32.0 g = 80.6 g = 80.6 g

Mass of the reactants equals mass of the products, — this further proves that not only atoms are conserved but also mass is conserved.

Using your balanced equations in Exercise 6, you can once again prove that mass is conserved. Review the previous example and work through the example below before going on to the next exercise. Consider Equation (e) from Exercise 6.

$$NaOH + H_2SO_4 \longrightarrow H_2O + Na_2SO_4$$

If the equation is balanced properly, the mass of the reactants should equal the mass of the products.

| Reactants | | Products | |
|---|---|---|---|
| 1 mol of NaOH | = 40.0 g | 1 mol of $H_2O$ | = 18.0 g |
| 1 mol of $H_2SO_4$ | = 98.1 g | 1 mol of $Na_2SO_4$ | = 142.1 g |
| total | 138.1 g | total | 160.1 g |

It is quite obvious that mass is not conserved. Balance the equation and prove that mass is conserved. Follow the outline below.

_____ NaOH + _____ $H_2SO_4$ ⟶ _____ $H_2O$ + _____ $Na_2SO_4$

| Reactants | Products |
|---|---|
| _____ mol NaOH = _____ g | _____ mol $H_2O$ = _____ g |
| _____ mol $H_2SO_4$ = _____ g | _____ mol $Na_2SO_4$ = _____ g |
| Total mass of Reactants = ____ | Total mass of products = ____ |

Are you convinced that mass is conserved? _____

Exercise 7

In similar manner, prove that mass is conserved, using balanced equations (h), (i), and (j), in Exercise 6.

Write the balanced equation first, then show your proof below.

1. Equation (h):

2. Equations (i):

3. Equation (j):

**Exercise 8**

In a balanced reaction:

1. Is it necessarily true that moles of atoms are always conserved? Explain.

_____

_____

_____

2. Given: $2HgO \longrightarrow 2Hg + O_2$
Is it necessarily true that moles of molecules are always conserved? Explain.

_____

_____

_____

## C. Using Equations in Problem Solving

Now that you have learned how to write and balance chemical equations you will be able to use chemical equations in problem solving. Since chemistry is a quantitative science, a chemist wishes to know more than the qualitative fact that a reaction occurs; he must also answer questions beginning: "How much ---" or "How many ---". The quantities are usually expressed in grams or moles as will be evident in the sample problems. In order to be successful in this type or work, a thorough understanding of the mole concept is essential. (Please review this concept if necessary). This quantitative relationship implied in a chemical reaction is called stoichiometry.

In this section, you will combine your knowledge on balancing equations, and what a mole is. Keep in mind that the coefficients in a balanced equation indicate the number of moles you have. Should you be given a quantity in grams, it must first be converted to moles before other quantities can be calculated. All of your coefficients will not be whole numbers. You could easily end up with fractions of a mole. The examples on page 12 will illustrate this.

Example 1

Suppose we wish to know how many moles of water are produced when 96.0 g of hydrazine ($N_2H_4$) are burned in a rocket.

The balanced equation for this reaction is:

$$N_2H_4 + O_2 \longrightarrow 2H_2O + N_2$$

The molar mass of hydrazine is 32.0 g/mol. In this problem 96.0 g of $N_2H_4$ are burned. How many moles of $N_2H_4$ is this?

$$\frac{96.0\ g}{32.0\ g/mol} = 3.0\ mol\ of\ N_2H_4$$

The balanced equation tells us that

$$1\ mol\ N_2H_4 \longrightarrow 2\ mol\ H_2O$$

This must mean that

$$3\ mol\ N_2H_4 \longrightarrow 6\ mol\ H_2O$$

We see that 96.0 g or 3 mol of hydrazine produce 6 mol of water. Since the molar mass of water is 18.0 g/mol, then 6 x 18.0 or 108 g of water would be formed.

You may use ratio and proportion to calculate moles of $H_2O$. This procedure may be handy when dealing with complex numbers.

$$\frac{1\ mol\ N_2H_4\ (from\ equation)}{3\ mol\ N_2H_4\ (given)} = \frac{2\ mol\ H_2O\ (from\ equation)}{X\ mol\ H_2O\ (required)}$$

Using numbers only we have

$$\frac{1}{3} = \frac{2}{X}$$ By cross multiplying

we obtain $X = 6$; $\therefore$ 6 mol of $H_2O$ are produced.

Example 2

When iron rusts it combines with oxygen in the air to form iron oxide, $Fe_2O_3$, according to the following balanced equation:

$$3O_2 + 4Fe \longrightarrow 2Fe_2O_3$$

Suppose we wish to calculate the mass of oxygen required for the complete reaction of 83.7 g of iron.

The molar mass of iron is 55.8 g/mol. In this problem 83.7 g of iron react. How many moles of iron react?

$$\frac{83.7\ g}{55.8\ g/mol} = 1.50\ mol\ of\ iron$$

The balanced equation tells us that

$$\underline{3}\ mol\ O_2\ react\ with\ \underline{4}\ mol\ Fe$$

This must mean that

$$3/4 \times 1.50\ mol\ or\ 1.12\ mol\ O_2\ react\ with\ 1.50\ mol\ Fe$$

To calculate moles of $O_2$ required, you may use ratio and proportion if you find it easier.

$$\frac{3\ mol\ O_2}{X\ mol\ O_2} = \frac{4\ mol\ Fe}{1.50\ mol\ Fe}$$

$$4X = 4.50$$
$$X = 1.12\ mol\ of\ O_2$$

Since the molar mass of oxygen is 32.0 g/mol, the mass of 1.12 mol of oxygen is:

$$1.12\ mol \times \frac{32.0\ g}{mol} = 35.8\ g$$

Therefore you may conclude that 35.8 g of oxygen will be required to react with 83.7 g of iron.

## Exercise 9

1. Methane, the main constituent of natural gas, has the formula $CH_4$. The products of combustion are carbon dioxide and water. Use spaces provided to answer questions (a) to (d).

   (a) Write the balanced equation for burning methane.

   (b) One mole of methane produces how many moles of water vapor?

(c) 0.125 mole of methane would produce how many moles of carbon dioxide?

(d) How many moles of water vapor would be produced by 4.00 g of methane?

If your answer to 1(d) was not "0.5", do Question 2 below.

2. Unbalanced: $Ch_4 + O_2 \longrightarrow CO_2 + H_2O$

Fill in the blanks with the correct coefficients.

(a) Balanced: _____ $CH_4$ + _____ $O_2 \longrightarrow$ _____ $CO_2$ + _____ $H_2O$

(b) Mole ratio: _____ mol + _____ mol $\longrightarrow$ _____ mol + _____ mol

(c) 0.125 mol $CH_4 \longrightarrow$ X mol $CO_2$

X = _____ mol of $CO_2$

Therefore 0.125 mol $CH_4$ produces _____ mol of $CO_2$.

(d) molar mass of $CH_4$ = 16.05 g/mol

Therefore 4.00 g of $CH_4 = \frac{g}{g/mol} = \frac{4.00\ g}{\_\_\_\_g/mol} = 0.25$ mol

__1__ mol of $CH_4$ produces _____ mol of water vapour (see equation (a)) 0.25 mol $CH_4$ (4.00 g) produces _____ mol water vapor.

## Exercise 10

1. Given $CaCO_3 \longrightarrow CaO + CO_2$.

(a) If 150.1 g of $CaCO_3$ are used in the above reaction, how many moles is this?

(b) How many moles of $CO_2$ would be produced by heating 150.1 g of $CaCO_3$?

(c) How many grams of $CO_2$ would be produced?

2. Given: $Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$

(a) If 4.0 g of hydrogen are produced by the above reaction, how many moles of hydrogen molecules are produced?

(b) How many grams of calcium would be required to produce 4.0 g of hydrogen?

3. Given that 48.6 g of magnesium, Mg, react with oxygen to yield a white solid, MgO:

(a) Write the balanced equation for this reaction.

(b) Calculate the number of moles of MgO that will be produced.

4. What mass of $Cl_2$ is needed to produce 63.5 g of $I_2$ when chlorine reacts with NaI according to the following equation (unbalanced):

$$NaI + Cl_2 \longrightarrow NaCl + I_2$$

5. Given: $Pb(NO_3)_2 + 2NaI \longrightarrow 2NaNO_3 + PbI_{2(s)}$

What mass of yellow precipitate, $PbI_2$, will be produced when 33.1 g of $Pb(NO_3)_2$ react. Show all steps below:

6. (a) If 71.0 g of chlorine gas and 10.0 g of hydrogen gas are allowed to react in a reaction chamber, calculate the mass of hydrogen chloride gas (HCl) produced.

(b) Which gas, if any, is left in excess, and state the amount.

## D. Types of Chemical Reactions

The different types of chemical reactions we will discuss are synthesis, decomposition, rearrangement, and ion-ion reactions. The name itself will give a pretty good idea as to what the reaction is like. However, the examples below will illustrate everything further.

### 1. Synthesis Reaction

When iron combines with sulphur upon heating, an entirely new product is formed, iron sulfide.

The equation for the above reaction is:

$$Fe + S \longrightarrow FeS$$

Another example is synthesis of ammonia, which is of great industrial use. Here, nitrogen reacts with hydrogen, under special conditions to produce ammonia ($NH_3$). The equation for this reaction is:

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

### Exercise 11

If 7.0 kg of nitrogen react, calculate the mass of hydrogen required for this reaction. 1 kg is 1000 g. The ratios are the same, whether you work in grams or kilograms.

## 2. Decomposition Reaction

You may have experienced a demonstration in other science courses where oxygen gas was produced by heating a compound such as potassium chlorate. This reaction shows decomposition of a compound into other simpler substances as shown by the equation below:

$$2KClO_3 \longrightarrow 2KCl + 3O_2$$

Probably you have also seen a demonstration on electroysis of water, where the compound water decomposes into its elements as shown by the following equation:

$$2H_2O \longrightarrow 2H_2 + O_2$$

## 3. Rearrangement or Replacement Reaction

In this type, two compounds react in such a way as to produce two new compounds. This reaction is evident when hydrochloric acid reacts with iron sulfide to produce an obnoxious smelling gas and another chemical residue. This reaction may be represented by the following equation:

$$FeS + 2HCl \longrightarrow H_2S + FeCl_2$$

## 4. Ion-ion Reaction

A yellow precipitate results when lead nitrate reacts with sodium iodide. This is a result of the reaction between the lead ions, $Pb^{+2}$, and the iodide ions, $I^-$, to produce lead iodide, $PbI_2$, a yellow precipitate.

Although the molecular equation does not show this, it is worthwhile to keep in mind the reacting ions which produce a new substance. The over all equation for the above reaction is:

$$Pb(NO_3)_2 + 2NaI \longrightarrow PbI_2 + 2NaNO_3$$

If everything that ionizes in water were to be written as an ion, and if everything that precipitates in water were to be written as a molecule, you would end up with the following ionic equation for the above reaction:

$$Pb^{+2} + 2NO_3^- + 2Na^+ + 2I^- \longrightarrow PbI_2 + 2Na^+ + 2NO_3^-$$

In the following exercise, assume molecules are involved unless ions are clearly indicated. Remember, if something does not have a (+) or (-) charge, it is not an ion.

Exercise 12 (Send for Correction)

Classify the following reactions as to type. Use the space provided.

(a) $S_8 + 8O_2 \longrightarrow 8SO_2$ _____

(b) $Al_2(SO_4)_3 + 3Ba(OH)_2 \longrightarrow 2Al(OH)_3 + 3BaSO_4$ _____

(c) $Na_2SO_4 + BaCl_2 \longrightarrow BaSO_4 + 2NaCl$
$2Na^+ + SO_4^{-2} + Ba^{+2} + 2Cl^- \longrightarrow BaSO_{4(s)} + 2Na^+ + 2Cl^-$ _____

(d) $Mg + ZnSO_4 \longrightarrow Zn + MgSO_4$
$Mg + Zn^{+2} + SO_4^{-2} \longrightarrow Zn + Mg^{+2} + SO_4^{-2}$ _____

(e) $H_2S \longrightarrow H_2 + S$ _____

(f) $Ca + Cl_2 \longrightarrow CaCl_2$ _____

(g) $N_2 + 6Li \longrightarrow 2Li_3N$ _____

(h) $2K + 2H_2O \longrightarrow H_2 + 2KOH$ _____

END OF LESSON E

# Chemistry Upgrading
# Lesson F

**DO NOT** send this lesson for correction

# CHEMICAL BONDING

A. Intra-molecular Bonding

1. Ionic Bonds
2. Covalent Bonds (Polar and Nonpolar)
3. Electron Dot Notation

B. Inter-molecular Bonding

1. Van der Waals Forces
2. Dipole-dipole
3. Hydrogen Bonding

## A. Intra-molecular Bonding

Most atoms do not exist as isolated atoms, but atoms are bonded together in one way or another. A bond is a force which tends to hold atoms together. As you know, atoms consist of electrons and protons. Since electrons and protons have opposite electrical charges, they attract each other. The protons are in the nucleus, and the electrons are outside the nucleus. A bond is formed when a single (outer) electron is attracted to two different nuclei at the same time. Two conditions must be met, however, before a bond will form. First of all, the resulting combination must have less energy than the individual atoms. Secondly, an atom tries to obtain an octet of electrons in its outer ring. Eight outer electrons seems to be a particularly stable arrangement, as is easily seen by how unreactive the noble gases are. Since hydrogen and lithium and beryllium have so few electrons, they only try to have two electrons in the outer ring.

There are two basic types of bond formation: electrons can be transferred from one atom to the next, or they can be shared between two atoms. What happens, will depend on the characteristics of the reacting atoms.

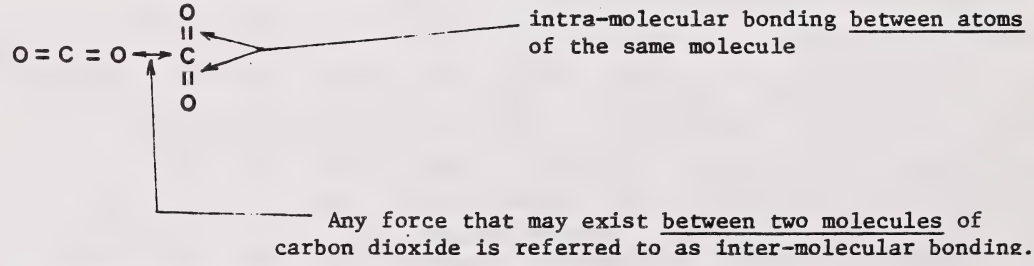
There are many differences between different atoms. Some atoms are smaller than others. Some atoms hold their outer electrons very tightly, while other atoms have a loose hold on their outer electrons. Some atoms must gain one or two electrons to obtain the electronic structure of the nearest inert gas. Other atoms must lose one or two electrons to have a stable outer octet of eight electrons. Some atoms have a great attraction for other electrons, while other atoms do not. Of course, atoms which need one or two electrons to complete their octet, have a high attraction for additional electrons. Therefore, atoms on the right side of the periodic table have a high electron affinity. The electron affinity is the energy released when an electron is added to a neutral atom. At the same time, the atoms on the right hold their own electron very tightly. Therefore they also have the highest ionization potential.

The ionization potential is the energy that is required to pull an electron off an isolated atom. An atom's placement on the periodic table will automatically give a pretty good indication as to what its ionization potential is. Generally speaking, the upper right hand corner will have the highest ionization potential, and the lower left hand corner will have the lowest ionization potential. This is because various trends can be noted in the periodic table, whether you move from left to right or from top to bottom, as was noted in Lesson C. It is not a coincidence that the atomic size is smallest in the upper right hand corner and largest in the lower left hand corner. On the extreme left, the alkali metals have a single electron in the outer shell, and this single electron is shielded from the nucleus by the inner electrons to some extent, so it tends to be far out, creating a large size.

However, you don t have to be concerned about all of these things when trying to figure out whether electrons are shared or transferred between two atoms. The electronegativity of each element takes these various factors into account, and the differences in electronegativity between two atoms will help you decide whether there is a transfer or a sharing of electrons. Electronegativity can be thought of as a combination between ionization potential and electron affinity. It varies from 4.0 in fluorine to 0.7 in francium. Electronegativity can be thought of as a relative attraction of an atom for electrons. Locate these elements on the periodic table. In general, the closer an element is to fluorine, the higher will be its electronegativity, and the closer an element is to francium, the lower will be its electronegativity. When two elements combine, the one with the higher electronegativity will have the greater pull on the shared or transferred electron. Inert gases do not have electronegativities since they don't react.

All of the above discussion enables one to determine different types of intra-molecular bonding.

Intra-molecular bonding refers to the type of bond that is formed between the atoms of a molecule. Inter-molecular bonding refers to forces of attraction between two molecules.


O
O=C=O
C
O
intra-molecular bonding between atoms of the same molecule
Any force that may exist between two molecules of carbon dioxide is referred to as inter-molecular bonding.

In virtually all types of bonding, however, we are mainly concerned with how the outer (negative) electrons of one atom are influenced by the (positive) nucleus of the next atom.

## 1. Ionic Bonds

In the previous section, it was seen that ionization energy, atomic size, electron affinity, and electronegativity determine what kind of bond you will have between two atoms. The most important factor is electronegativity. The attraction an atom has for its own outer electrons is its ionization potential. The attraction an atom has for the next atom's outer electron is its electron affinity. However, only differences in electronegativity need to be considered when determining whether or not a compound is ionic since electronegativity is a combination of ionization energy and electron affinity. As with many things, there is a gray area as to where the borderline is between ionic bonds and bonds which are not ionic, just like there is no precise answer as to the exact minute that the headlights of a car should be turned on at night. Arbitrarily, assume that when the difference of electronegativity is 1.8 or more, that you have an ionic bond. An ionic bond occurs when the electrons are considered to be completely transferred to the more electronegative atom. Then you would essentially have a positive ion attracted to a negative ion. This is an ionic bond. A table of electronegativities is found on the periodic table of Lesson B.

Generally speaking, when two atoms from opposite ends of the periodic table combine, an ionic bond will be formed.

## 2. Covalent Bonds

In covalent bonds, electrons are shared, although they are not necessarily shared equally.

Covalent bonds can be polar or nonpolar, but in all covalent bonds, electrons are shared, at least to some degree. In ionic bonds, electrons are considered to be completely transferred. If you have no difference in electronegativity between two atoms, you have a nonpolar covalent bond. If the difference in electronegativity is between 0.1 and 1.7, you have a polar covalent bond. If the difference in electronegativity is 1.8 or greater, you have an ionic bond.

Nonpolar covalent bonds have no positive or negative poles. The electrons are shared equally since no atom has the greater attraction for electrons. At the other extreme, we have ionic bonds, in which electrons are completely transferred. In between, there are polar covalent bonds. In some cases, the bond is almost ionic, and at the other end, one atom only has a slightly greater pull on the electrons.

| ionic | polar covalent | nonpolar covalent |
|---|---|---|
| complete transfer of electrons | unequal sharing of electrons<br>i.e. poles are present | equal sharing of electrons<br>i.e. no poles. |

In an ionic bond like NaCl, the Cl end has a charge of -1, and the Na end has a charge of +1. In a nonpolar bond like $Cl_2$, neither end has any charge. In a polar covalent compound like AlP, the phosphorus atom exerts a greater pull on the electrons which are shared between the two atoms, and you might think of the aluminum end as having a charge of $+\frac{1}{10}$ with the phosphorus end having a charge of $-\frac{1}{10}$. The bonding electrons are only <u>partially</u> transferred to phosphorus, so the bonding is polar covalent and not ionic.

## Exercise 1

1. Define the following terms:

(a) chemical bond _____

_____

_____

(b) ionization energy _____

_____

_____

(c) electronegativity _____

_____

_____

(d) ionic bond _____

_____

_____

(e) polar covalent bond _____

_____

_____

(f) non polar covalent bond _____

_____

_____

2. State the requirements for a stable chemical bond.

_____

_____

_____

3. Which single property is generally used to determine if a bond is nonpolar, polar, or ionic?

_____

4. For the following compounds, state whether the bond is nonpolar covalent, polar covalent, or ionic: KCl, $H_2$, $H_2O$, $I_2$, BrCl, ZnS, $CaCl_2$, $Al_2O_3$, AgCl

| nonpolar covalent | polar covalent | ionic |
|---|---|---|
| | | |
| | | |
| | | |
| | | |
| | | |
| | | |

5. (a) Which corner of the periodic table has the highest electronegativity?

_____

(b) Which corner of the periodic table has the lowest electronegativity?

_____

6. Which element has the greater attraction for electrons, and would therefore be at least partially negatively charged in the following compounds?

BeS **S**, BrCl _____, MgO _____, $Cl_2O$ _____,

$OF_2$ _____.

7. Which element is always negatively charged when it forms a compound with another element? _____

8. Underline the correct word below.

   (a) Across a period, atomic size (increases or decreases).

   (b) Across a period, ionization energy (increases or decreases).

   (c) Across a period, electronegativity (increases or decreases).

   (d) Down a group, atomic size (increases or decreases).

   (e) Down a group, ionization energy (increases or decreases).

   (f) Down a group, electronegativity (increases or decreases).

   (g) The smaller the atom, the (smaller or larger) is the electronegativity.

   (h) The smaller the atom, the (smaller or larger) is the ionization energy.

   (i) The smaller atom will be (positively or negatively) charged in a compound.

9. Describe the bonding of two hydrogen atoms to form one molecule of hydrogen gas ($H_2$) in terms of energy. _____

_____

_____

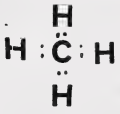
3. Electron Dot Notation

The electron dot notation is a way of showing all outer electrons in a compound. Note that outer electrons are also called valence electrons. Each dot represents one outer electron. If two or four dots are shown between two atoms, this means that two or four electrons are shared between the two corresponding atoms. Thus a bond is created because certain electrons are attracted to two nuclei at the same time.

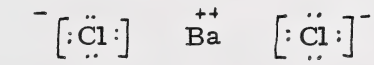
Besides the attraction of opposite charges, there are other rules which must be followed when writing electron dot structures of compounds.

(1) The group number of an element determines how many dots are to surround each element. For example, calcium is in group IIA, so the electron dot configuration for calcium is Ca: and since sulfur is in group VIA, its electron dot configuration is $\dot{\underset{\cdot\cdot}{\mathrm{S}}}:$ .

(2) When elements combine, they would like to have an octet of 8 electrons in their outer shell. This can be accomplished by either sharing additional electrons from the next atom, which happens with covalent bonding, or by transferring the necessary electrons from one element to the next, which happens in the case of ionic bonding. To illustrate how this octet is arrived at with covalent bonding, consider $CH_4$. Since carbon is in group IVA, its electron dot structure is $\cdot\dot{C}\cdot$, and since hydrogen is in group IA, its structure is $H\cdot$. Since carbon needs four additional electrons to get an octet, you need four hydrogens, which results in the following electron dot structures for $CH_4$:



Hydrogen is one of the few elements, along with Li, Be and possibly B, which only needs 2 electrons instead of 8 for a complete outer shell. To illustrate how an octet is arrived at with ionic bonding, consider $BaCl_2$. Ba is in group IIA, which gives $Ba:$, and Cl is in group VIIA, which gives $:\dot{\underset{\cdot\cdot}{Cl}}\cdot$, so the structure for $BaCl_2$ is:



With ionic bonding, count the number of electrons that the negative ion obtained, and put the appropriate charge on that ion. Since each chlorine gained one electron, its charge is -1. Similarly, count the number of electrons lost by the positive ion, and put the appropriate charge in that ion. In the above case, Ba lost 2 electrons, so its charge is +2. Note that chlorine now has an octet of eight electrons. When a metal loses all of its outer electrons, assume that what is now the outer shell of the ion has 8 electrons, although no electrons are shown since the valence electrons are lost.

(3) When writing electron dot structures, always put the electrons in pairs, between the various atoms. Usually, there is only one pair of electrons between atoms, as in the case of BrCl, $:\dot{\underset{\cdot\cdot}{Br}}:\dot{\underset{\cdot\cdot}{Cl}}:$. However, it is possible to have 2 pairs of electrons between atoms, as in the case of $C_2H_4$,

Note that this electron dot structure follows the previous rules. Each carbon atom has an octet of eight outer electrons in each circle. Also, since carbon is in group IVA, and hydrogen is in group IA, the total number

of outer electrons for $C_2H_4$ should be 4 + 4 + 1 + 1 + 1 + 1 = 12. $C_2H_4$ has a double bond, but triple bonds are also possible.

Exercise 2

Write the electron dot structure for the following molecules:

ClF

LiF

$CCl_3Br$

$C_2H_2$

$C_2H_4$

$C_2H_6$

Optional: If you want a real challenge, try $CH_3COOH$

B. Inter-molecular Bonding

Inter-molecular Bonding refers to bonding between two different molecules, and NOT within a molecule. Different types of inter-molecular bonding include van der Waals forces, dipole-dipole forces, and hydrogen bonding.

1. Van der Waals Forces.

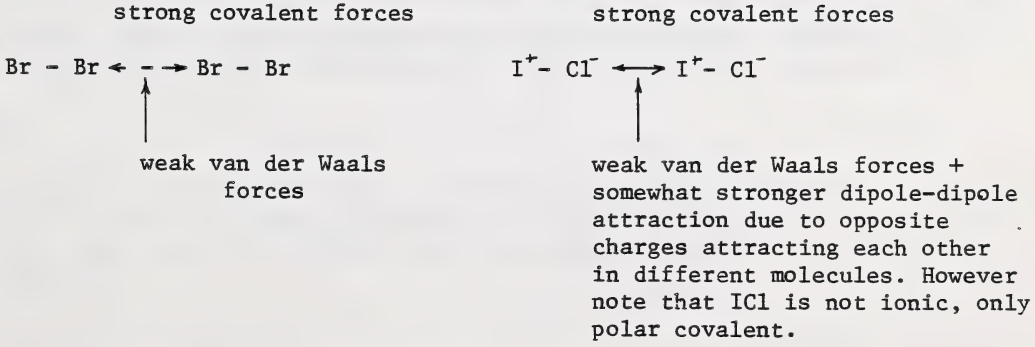
Van der Waals forces are extremely weak forces which are present between different molecules. Each molecule contains a large number of electrons, and these electrons are constantly in motion. It is possible that some of these electrons may be more toward one side of the molecule than at other times. As a result, the side with the greater number of electrons will be slightly negative for a split second, and the side with fewer electrons will be positive for a split

second. This has a slight effect on neighboring electrons and molecules, and as a result, various small attractions between molecules occur. These very weak attractive forces are called van der Waals forces. Usually, other forces of attraction are also present between molecules, and these other forces are usually much greater than van der Waals forces.

2. Dipole-dipole forces

There are several forces which act simultaneously between two molecules and it is not easy to just isolate one factor and say that everything is due to one thing. For example, van der Waals forces exist between two molecules. Everything else being equal, the heavier that the molecule is, the higher the boiling point will be. Dipole-dipole forces also exist between two molecules, and a higher melting point may be due to a combination of dipole-dipole forces and van der Waals forces.

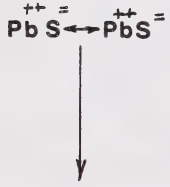
A dipole-dipole force can only exist between two molecules which are either polar covalent or ionic. A dipole-dipole force results due to a positive to negative attraction between molecules. For example, compare ICl and $Br_2$. $Br_2$ is nonpolar and its melting point is -7.2°C and its boiling point is 59°C. One bromine molecule is attracted to the next bromine molecule by van der Waals forces. ICl is polar and its melting point is 27°C and its boiling point is 97°C. Both the melting point and boiling point for ICl is higher than for $Br_2$. Both ICl and $Br_2$ have the same number of electrons, therefore the van der Waals forces are comparable. ICl has additional dipole-dipole forces which results in a higher temperature being required to melt and boil it. Below, is a diagram comparing forces in $Br_2$ and ICl.


strong covalent forces
Br - Br ← - → Br - Br
weak van der Waals forces
strong covalent forces
I+ - Cl- ⟷ I+ - Cl-
weak van der Waals forces + somewhat stronger dipole-dipole attraction due to opposite charges attracting each other in different molecules. However note that ICl is not ionic, only polar covalent.

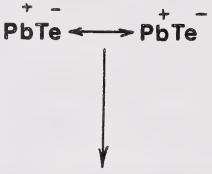
Further, compare PbS with PbTe. The electronegativity of Pb is 1.8, that of S is 2.5, and that of Te is 2.1. This clearly shows that PbS and PbTe are polar covalent. In other words, the Pb end of the molecule is slightly positive, and the S or Te end is

slightly negative. The PbS molecule can be represented like this + - . Now, what would happen when several PbS molecules are close together? The Pb end of one molecule would be attracted to the S end of the next molecule. This attraction would be much weaker than the attraction within a molecule, although the attraction is stronger than that between nonpolar molecules.

The melting point of PbS is 1120°C and the melting point of PbTe is 917°C, despite the fact that PbTe has higher van der Waals forces due to its greater number of electrons. The reason for the higher melting point is due to larger dipole-dipole forces between PbS. There is a greater difference in electronegativity between Pb and S than between Pb and Te. As a result, the Pb would be more positive in PbS, so there is a greater attraction between the Pb atom of one molecule and the S atom of the next molecule.




larger dipole attraction




smaller dipole attraction

PbS and PbTe are covalently bonded, so you can't talk of ions. But you can say that Pb is partially positively charged and Te is partially negatively charged. For example Pb might be + $\frac{1}{8}$ and Te might be - $\frac{1}{8}$ instead of +1 and -1 as in the case of ions. With PbS, Pb might be + $\frac{1}{5}$ and S might be - $\frac{1}{5}$, hence the larger dipole-dipole attraction.

Now, consider an ionic substance like NaCl. There are no molecules as such in NaCl. You have an aggregate of ions.

eg.

$Na^+ Cl^- Na^+ Cl^-$

$Cl^- Na^+ Cl^- Na^+$

$Na^+ Cl^- (Na^+) Cl^-$

$Cl^- Na^+ Cl^- Na^+$

The circled $Na^+$ is equally attracted to $6Cl^-$ ions (also front and back in 3 dimensions). In PbTe, one Pb was completely attracted to one Te, and there was a partial attraction to the Te of the next molecule. This partial attraction was the dipole-dipole force of attraction between 2 molecules. With something like NaCl, you might say that the dipole-dipole force is as strong as it can possibley get, since each $Na^+$ is attracted equally strongly to all

neighboring $Cl^-$. The more ionic a substance is, the greater is the dipole-dipole force. The closer a substance is to having nonpolar covalent bonds, the weaker is the dipole-dipole force of attraction between two different molecules.

3. Hydrogen Bonding

Hydrogen bonds are much like dipole-dipole forces. The difference is that hydrogen has to be one of the polar molecules. This hydrogen is usually attracted to the negative end of the neighboring molecule. Hydrogen bonds are most important when there is a large difference in electronegativity between hydrogen and the other element. That is why hydrogen bonds form with F, O, and N. In other words, the hydrogen has to be covalently bonded to an F, O, or N in the molecule it is in, and also it must be hydrogen bonded to an F, O, or N of the next molecule. Due to the high electronegativity and small size of F, O, and N, you almost have a single proton between two of these highly electronegative atoms, hence the large inter-molecular bond. To illustrate the effect hydrogen bonds have, consider some boiling points. $H_2O$ boils at 100°C, $H_2S$ boils at -59.6°C, despite the fact that $H_2S$ is heavier and would therefore normally have a higher boiling point than water. Thus we see that hydrogen bonds make it more difficult fo one water molecule to break away from the next water molecule. Hydrogen bonds are about one-tenth as strong as the bond between hydrogen and oxygen atoms of the same molecule. In turn, hydrogen bonds are ten times stronger than the bond between two different nonpolar molecules such as $O_2$.

Everything else being equal, hydrogen bonds are stronger than dipole-dipole forces, which in turn are stronger than van der Waals forces. The more polar a molecule is, the stronger the dipole-dipole forces are, with ionic bonds having the strongest dipole-dipole forces. The stronger the intermolecular bonds are, the higher the melting and boiling points are.

Exercise 3

1. $PbF_2$ melts at 855°C and boils at 1290°C. $PbI_2$ melts at 402°C and boils at 954°C. Why? (In your explanation discuss the types of forces involved and the atoms they act on and state which forces are stronger and which are weaker.)

_____

_____

_____

_____

2. Which would you expect to have the higher melting and boiling point: $Cl_2$ or $O_2$? _____ Why? _____

3. Which would you expect to have the higher melting and boiling point: $O_2$ or NO? _____ Why? _____

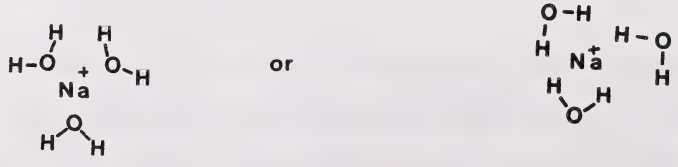
4. Which would have the higher boiling point, HF or HBr? _____

   Why? _____

5. Various ionic substances readily dissolve in water.

   (a) From the shape of the water molecule and from what you have learned about electronegativities and bonding, which picture accurately shows an $Na^+$ dissolved in water?

   (i) (ii)



5. (b) Which picture accurately describes a $Cl^-$ ion dissolved in water.

(i) (ii)

6. If any substance can be smelled, this implies that molecules in one form or another are being given off by the substance, which in turn indicates relatively weak inter-molecular bonds. True or False?

7. Every chemical which gives off an odor will get smaller and eventually will "disappear" after enough time has elapsed. True or False?

8. Using the electronegativity table, predict whether each of the following substances is ionic or molecular. Molecular substances are covalent.

(a) sulfur _____

(b) calcium chloride _____

(c) iodine _____

(d) sulfur dioxide _____

(e) any hydrocarbon _____

9. Which of the above substances would have a higher melting point than the rest?

_____

END OF LESSON F

# Chemistry Upgrading
# Lesson G

**DO NOT** send this lesson for correction

# SOLUTIONS

A. Solutions are Mixtures

1. Homogeneous Mixtures of Solute and Solvent
2. Classification of Solutions
   (a) Gaseous Solutions
   (b) Liquid Solutions
   (c) Solid Solutions

B. The Dissolving Process of Solutions

1. Ions in Solution
2. Molecules in Solution
3. Energy Changes
4. Electrolytes and Nonelectrolytes

C. Concentration of Solutions

1. A Molar Solution
2. A $mol/m^3$ Solution
3. Quantitative Aspects
   (a) Preparation of Solutions
   (b) Determining Molarity of Solutions
   (c) Using Molarity of Solutions in Problem Solving
   (d) Dilution of Solutions

## A. Solutions are Mixtures

### 1. Homogeneous Mixtures of Solute and Solvent

When sand and water are mixed in a beaker the sand settles to the bottom of the beaker. There are distinct parts to the mixture. The word HETEROGENEOUS is used to describe this system. Heterogeneous mixtures consist of distinct phases whose properties remain more or less defined.

When a small amount of sugar and water are mixed, the sugar dissolves completely. The solid disappears to become part of the liquid. Mixtures of this type are called SOLUTIONS. The word HOMOGENEOUS is used to describe such a system. Both pure substances and solutions are homogeneous. In other words, homogeneous mixtures consist of a single phase whose properties may be considerably different from those of its components.

Whenever a solid dissolves in a liquid, the solid is called the solute and the liquid is called the solvent. Thus in a salt water solution, the salt is the solute and the water is the solvent.

With water solutions, the solute particles are uniformly distributed; they are molecular in size; they cannot be removed by filtering; they do not settle out; and they form a transparent mixture.

Exercise 1

1. List five properties of water solutions.

   (a) _____

   (b) _____

   (c) _____

   (d) _____

   (e) _____

2. Fill in the blanks with the appropriate term.

   (a) In liquid solutions the substance dissolved is called the _____.

   (b) The substance dissolving the solid (like water dissolving sugar) is called the _____.

## 2. Classification of Solutions

There are three types or physical states of solutions:

(a) Gaseous Solutions
(b) Liquid Solutions
(c) Solid Solutions

Solutions may be composed of any combination of the three physical states hence nine classes of solutions are possible.

### (a) Gaseous Solutions

All gas mixtures are homogeneous, therefore gas mixtures are solutions and appear in only one phase – the gas phase. The molecules behave as gas molecules, although these molecules

may have come from gaseous, liquid or solid substances. Examples of the latter two include perfume in air and carbon in the form of smoke.

Whatever the source of the components, this gaseous solution is a single, homogeneous phase.

You will note that air is a good example of a gaseous solution. As with other solutions, components of air can be separated by phase changes. Through a low-temperature distillation process, liquid air can be separated into liquid oxygen, liquid nitrogen and other substances. Below is a table showing the composition of dry air.

Composition of Dry Air

| Name | Formula | Percentage of Molecules |
|---|---|---|
| Nitrogen | $N_2$ | 78.06 |
| Oxygen | $O_2$ | 20.99 |
| Argon | Ar | 0.93 |
| Carbon dioxide | $CO_2$ | 0.03 |

Trace quantities of hydrogen, krypton, xenon, and neon are also present in the air. Their total contribution is less than 0.003%. The air that we breathe is never really dry, however, and may have up to 2% water vapor.

(b) Liquid Solutions

Liquid solutions can be made by mixing two liquids,(e.g. lemon juice and water) by dissolving a gas in a liquid,(e.g. oxygen in water supports marine life) or by dissolving a solid in a liquid (e.g. sugar in water). The result is a homogeneous system containing more than one substance. By definition, the result is a solution.

(c) Solid Solutions

Solid solutions are very common in nature. Many of our common minerals and gem stones such as emeralds, sapphires, and rubies are solid solutions which have fascinated people over many ages. Gold atoms can replace some of the copper atoms in a copper crystal to give a gold-copper alloy. The alloy is a solid solution. In the same manner, copper atoms can replace gold atoms in a gold crystal to give another solid solution, a copper-gold alloy. Many other alloys are solid solutions.

Solid solutions are very useful today in transistors, amplifiers, record players and television units.

Another example of a solid solution is when mercury, a liquid, dissolves in gold. Also, hydrogen, a gas, can dissolve in solid palladium.

## Exercise 2

Classify each of the following solutions. After each example fill in the blanks under the appropriate headings. Two examples have been done for you.

| Example | State of Solute | Name of Solute | State of Solvent | Name of Solvent | Class |
|---|---|---|---|---|---|
| A bottle of carbonated soft drink | | | | | |
| Lemonade | | | | | |
| Antifreeze in your car (30% ethylene glycol) | liquid | ethylene glycol | liquid | water | liquid-liquid |
| Gas used in certain medical situations e.g. respiratory restrictions | gas | oxygen | gas | helium | gas-gas |

## B. The Dissolving Process of Solutions

### 1. Ions in Solution

To review briefly the nature of an ion, you must keep in mind that an atom is electrically neutral, that means that the number of protons (positively charged particles) always equals the number of electrons (negatively charged particles) in a neutral atom or molecule. We also say that they have zero electric charge.

Certain atoms have a tendency to gain electrons while others have a tendency to lose electrons. When these occurences take place, particles are formed with different numbers of electrons and protons. There is a net electric charge on these particles. Chemists have named these charged particles IONS. In other words an ion can have more electrons than protons, or more protons than electrons.

To illustrate this further, consider a chlorine atom. As a neutral atom it has 17 electrons (symbol $e^-$) and 17 protons (symbol $p^+$). A chlorine atom and its related family has a tendency to pick up an electron therefore giving it a net charge of -1.

$$Cl + e^- \longrightarrow Cl^-$$

$$\begin{matrix} 17e^- \\ 17p^+ \end{matrix} + e^- \longrightarrow \begin{matrix} 18e^- \\ 17p^+ \end{matrix}$$

You will note that the electron population changes from 17 to 18 which gives the atom an over-all or net charge of -1.

This addition of an electron to a neutral atom produces a charged particle called a chloride ion ($Cl^-$).

To check further, we have:

17 protons + 18 electrons $\longrightarrow$ net charge of negative one

$$(17^+) + (18^-) \longrightarrow (-1)$$

Take another example, where a neutral atom loses an electron to become a positively charged particle. Potassium and its related family of elements is a good example of this behavior. Here again we note that a potassium atom has 19 electrons and 19 protons. (Should you wonder how we obtain this number, consult with your periodic table and note that the atomic number represents the proton number which also equals the electron number in a neutral atom).

In symbol form we can represent a potassium atom losing an electron to become a potassium ion, with a positive charge as follows:

$$K \longrightarrow K^{+} + e^{-}$$

$$\begin{matrix} 19e^{-} \\ 19p^{+} \end{matrix} \longrightarrow \begin{matrix} 18e^{-} \\ 19p^{+} \end{matrix} + e^{-}$$

You will note that the electron population changes from 19 to 18 which gives the atom an over-all or net charge of +1. Thus the removal of an electron from a neutral potassium atom produces a charged particle called a potassium ion. ($K^{+}$)

To check further we have:

19 protons + 18 electrons $\longrightarrow$ net charge of positive one

$(19^{+})$ + $(18^{-})$ $\longrightarrow$ +1

By using all this knowledge about how ions are formed, we can go further to explain the dissolving process of compounds in water. For example, take the common compound, sodium chloride (NaCl) and see what happens to it when it dissolves in water. Here it can be found experimentally that sodium chloride will break up or dissociate into two distinctly charged particles, positive and negative particles. Can you guess what these particles are? If not, we can depend on our previous theory to explain this phenomenon. When sodium chloride is placed into water, by means of a complex dissolving process the forces which bind these particles in a salt crystal weaken, and will cause these particles to go into solution as charged particles called ions. The two types of ions formed are sodium ($Na^{+}$), having a +1 charge and chloride ($Cl^{-}$) having a -1 charge. This process of dissociation of compounds in water is known as ionization of compounds.

The positively charged ions are generally called CATIONS while the negatively charged ions are called ANIONS.

A list of common ions appears in your periodic table. You will also note that charges on various ions differ. This is because certain atoms may undergo a change of one, two or more electrons. You will also note that the sign for the charge follows the numerical value. e.g. $SO_4^{2-}$ , $Cl^{1-}$ . It is also correct to use the sign preceding the number as in $SO_4^{-2}$, $Cl^{-1}$.

Always treat all radicals or complex ions as a unit.

Exercise 3

Show by using an equation how each of the following compounds form ions when in solution. Two examples have been worked out for you. (Refer to the periodic table of ions if necessary).

1. $AgNO_3 \longrightarrow Ag^+ + NO_3^-$
2. $Al_2(SO_4)_3 \longrightarrow 2Al^{+3} + 3SO_4^{-2}$
3. $KI \longrightarrow$
4. $CuSO_4 \longrightarrow$
5. $MgCl_2 \longrightarrow$
6. $Na_2S_2O_3 \longrightarrow$
7. $Fe_2(SO_4)_3 \longrightarrow$
8. $NH_4Cl \longrightarrow$
9. $ZnCO_3 \longrightarrow$
10. $Ba(NO_3)_2 \longrightarrow$

NOTE: The total number of positive charges must equal the total number of negative charges.

2. Molecules in Solution

Compounds may also dissolve in liquids but remain in molecular form. That means that they do not ionize and no charged particles or ions appear in the resulting solution. Sugar dissolves in water but does not produce any ions.

Sugar (solid) + $H_2O \longrightarrow$ Sugar (solution)

or $C_{12}H_{22}O_{11\,(s)} + H_2O_{(l)} \longrightarrow C_{12}H_{22}O_{11\,(aq)}$

A method used for determining which compounds will form ions and which will not will be discussed later in this lesson under the topic: electrolytes and nonelectrolytes.

Some substances such as vinegar will produce solutions which may be composed of a mixture of both ions and molecules. It means that some molecules of vinegar will dissociate into ions, while other molecules of vinegar will remain as molecules in the

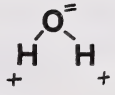
water solution. The degree of dissociation depends upon such factors as: nature of solute, nature of solvent, dilution, and the temperature at which dissolving takes place.

3. Energy Changes

When compounds dissolve in water, changes in the temperature of the resulting solution may be evident. That is, certain solutions get warmer while others get colder during the dissolving process. How can these temperature changes be explained? Let us consider the dissolving of sodium chloride in water to illustrate some aspects of the dissolving process. Within the solid salt crystal there are electrical interactions or binding forces holding the positive sodium ions and the negative chloride ions. In the solid, these ions are close together and the compound is in its lowest state of potential energy. Upon dissolving the compound, energy is added to separate these ions. Now the potential energy increases since distance between positive and negative ions icreases.

Other energy changes take place when the compound is placed into water. Strong electrical interactions occur between water molecules and the ions of the salt in solution. To understand these electrical interactions you must know more about the nature of the water molecule. We say that a water molecule has an electric dipole or that it is a polar molecule. What does this mean? To explain electric dipoles it is necessary to study the symmetry of the molecule which in turn will give us a better picture of how the electrons are distributed around the nuclei of atoms. To explain symmetry you must consider the geometrical shape of a molecule. You probably already know that some molecules are bent, while others are linear. Some form pyramidal shapes while others form tetrahedral shapes. All these different geometric shapes affect the charge distribution within the molecule. Since our concern is primarily about the water molecule, you will see the bent structure of the molecule from the diagrams below.

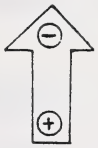
or




$H_2O$

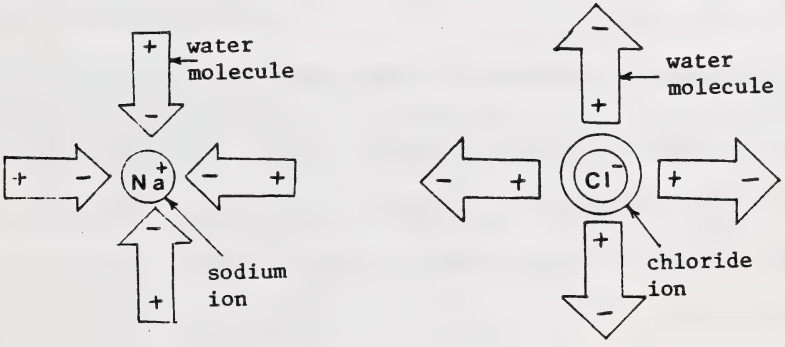
In the above diagrams note that two hydrogen atoms are bonded to one oxygen atom. Since the oxygen atom has a greater attraction for electrons than the two hydrogen atoms, an unequal charge distribution occurs since there will be a "spilling over" of electrons toward the oxygen end of the molecule, giving the molecule a negative character on one end and a positive character on the other as indicated in the diagrams above. This unequal distribution of

electrons within a molecule gives it a polar character, also known as a dipole. This molecular dipole can be represented by an arrow. The water molecule is not completely ionic but very polar covalent, so the water molecule would not split apart from form $2H^+$ and $O^{-2}$, as is the case with an ionic substance like NaCl.



Going back to our original problem, how then does dissolving sodium chloride in water affect energy?

When an electric dipole, as represented by a water molecule, is brought near an ion, the polar water molecule surrounds the ion in such a way that the negative end of the water molecule is nearest the positive sodium ion while the positive end of the water molecule is nearer the negative chloride ion (similar to N and S magnetic interactions). When this particular orientation takes place, we say that HYDRATION has occurred. (See diagram below.)




But what about the energy changes? When sodium and chloride ions become surrounded by water molecules, unlike charges are close together, and now potential energy decreases. When this happens, energy is released and this energy is called hydration energy.

From all this discussion you may now want to know what is the net effect on energy associated with dissolving sodium chloride in water. Experimentally it can be determined whether more energy was absorbed in separating the ions, which is the lattice energy, or whether more energy was released during hydration. If more energy is absorbed than released, the temperature of the water (solution) decreases, then this reaction is called an endothermic reaction. If more energy is released than is absorbed then the temperature of the water increases and then this reaction is called an exothermic reaction.

At this point we will not worry about the actual amounts of energy involved in the dissolving process.

In summary, the following three things happen when something dissolves.

1. Solute particles separate;
2. Solvent particles move apart, providing space for the solute particles;
3. Solvent and solute particles attract one another.

On dissolving, the temperature may go up or down or remain virtually unchanged, depending on the relationship between lattice energy and hydration energy.

Exercise 4

1. Everything that dissovles will automatically produce ions. True or false? _____

2. Why is the water molecule a polar molecule? _____

_____

_____

3. Towards which end of the water molecule will a positive ion be close to? _____ Why? _____

_____

4. The hydration energy is exothermic. True or false? _____

5. The lattice energy is endothermic. True or false? _____

## 4. Electrolytes and Nonelectrolytes

Solutions are classified by their ability to conduct electricity. It was Michael Faraday who first did extensive study and research on electrical conductivity of solutions and published a report on his findings in 1833. He came to the conclusion that charged particles in solution were responsible for the conductivity of electricity. He named these conducting solutions ELECTROLYTES. It was later that Svante Arrhenius (1850-1927) published a report on the properties of electrolytes. He believed that certain substances when dissolved

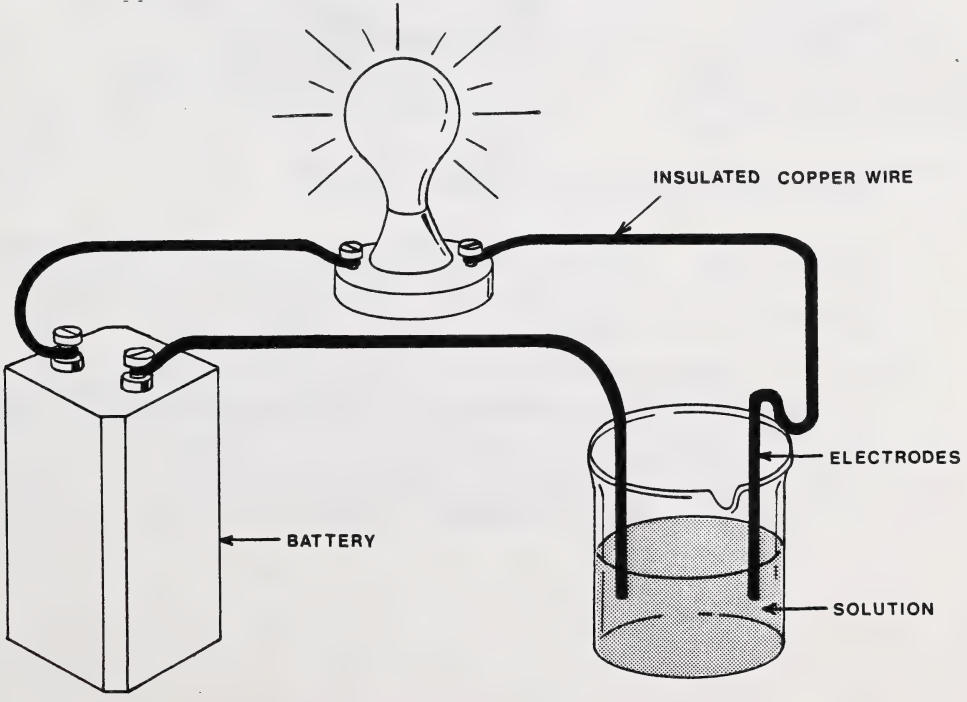
in water produced both positive and negative ions. His research in this area brought about the Theory of Ionization, which is used by chemists today.

Solutions, which do not conduct electricity are classified as NONELECTROLYTES. A nonelectrolyte is defined as a substance which does not form ions in solution; almost all of it remains in molecular form. An electrolyte may have both molecules and ions in solution. If there are enough ions to conduct a strong current, the solution is called a strong electrolyte. On the other hand, solutions may contain small concentrations of ions and will conduct a current weakly. These solutions are called weak electrolytes. Some solutions are so weak that for all practical purposes we can classify them as nonelectrolytes.

An example of a strong electrolyte is sodium chloride in water. Here, the presence of large amounts of sodium and chloride ions make the solution a strong electrolyte. Sugar dissolved in water will not conduct a current because sugar does not form ions but dissolves to form molecules of sugar. Therefore a sugar solution is called a nonelectrolyte.

You may wonder at this point how we can determine which solutions are strong, weak, or nonelectrolytes. Apparatus like that shown below may be used to show conductivity of solutions.

Apparatus to Show Conductivity of Solutions



If a solution is a strong electrolyte, the electrons will travel through the circuit to cause a bright glow on the bulb. If a solution is a nonelectrolyte or a weak electrolyte the bulb will not light up, or if it does, it will only glow very weakly when the circuit is set up as shown in the diagram above.

## Exercise 5

1. What is a characteristic of a strong electrolyte? _____

_____

2. Which compounds would most likely form strong electrolytes, those with a small difference in electronegativity, or those with a large difference in electronegativity?

_____

## C. Concentration of Solutions

1. The properties of solutions vary dependent upon the amount of solute to the amount of solvent. The concentration of a solution is expressed as the number of moles of solute in one litre of solvent. A formula is often used in solving problems using concentrations of solutions.

$$C = \frac{n}{V}$$ where C = concentration
n = number of moles of solute
V = volume of solvent

2. Preparation of Solutions

A student is required to prepare 1L of a 1 mol/L solution of sodium chloride (NaCl). Standard laboratory equipment includes a 1 L volumetric flask (a flask with a long narrow neck that is etched with a line indicating a 1 L volume).

Step 1: Determine the number of moles of sodium chloride required. From the formula given, rearrangement produces the formula.

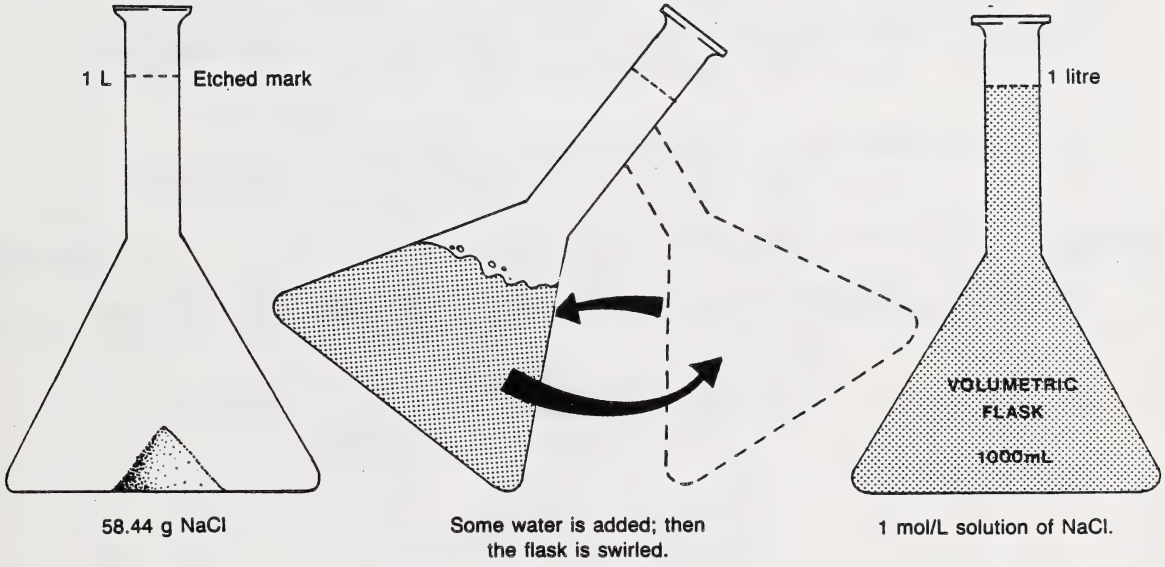
$$n = CV$$
$$n = (1 \text{ mol/L})(1 \text{ L})$$
$$n = 1 \text{ mol}$$

Step 2: Determine the mass of 1 mol of NaCl. From the periodic table:

1 mol Na = 22.99 g
1 mol Cl = 35.45 g
1 mol NaCl = 58.44 g

Step 3: The student uses a balance to obtain 58.44 g of NaCl.

Step 4: Pour the NaCl into a volumetric flask and add some water. Swirl to dissolve the salt then add water up to the 1 L mark. This is a 1 mol/L solution.




## Examples

1. Explain how you would prepare 500 mL of a 0.250 mol/L solution of KCl.

Answer: Let us first assume we have only the 1 L volumetric flask. This means we will have to prepare 1 L of solution and then measure out 500 mL using a graduated cylinder. The remaining 500 mL will have to be discarded or stored. (If a 500 mL volumetric flask is available see case II of this answer).

Step 1: Calculate the number of moles of KCl required:

$$n = CV$$
$$= (0.250 \text{ mol/L})(1 \text{ L})$$
$$= 0.250 \text{ mol}$$

Step 2: Calculate the mass of 0.250 mol of KCl.

$$m = nM$$
$$= (0.250 \text{ mol})(74.55 \text{ g/mol})$$
$$= 18.6 \text{ g}$$

Step 3: Use a balance to obtain 18.6 g of KCl

Step 4: Pour the KCl into a 1 L volumetric flask, add some water, swirl until dissolved, and add water to the graduated mark.

Step 5: Measure out 500 mL of this 0.250 mol/L solution.

Case II If a 500 mL volumetric flask is available:

Step 1: Calculate the number of moles of KCl required:

$$n = CV$$
$$= (0.250 \text{ mol/L})(0.500 \text{ L})$$
$$= 0.125 \text{ mol}$$

Step 2:

$$m = nM$$
$$= (0.125 \text{ mol})(74.55 \text{ g/mol})$$
$$= 9.32 \text{ g}$$

Step 3: Use a balance to obtain 9.32 g of KCl.

Step 4: Pour the KCl into a 500 mL volumetric flask, add some water, swirl to dissolve, add water to the 0.500 L mark.

2. How many grams of $PbCl_2$ are required to make 1.000 L of a 0.350 mol/L solution?

$$n = CV$$
$$= (0.350 \text{ mol/L})(1.000 \text{ L})$$
$$= 0.350 \text{ mol}$$

$$m = nM$$
$$= (0.350 \text{ mol})(278.05 \text{ g/mol})$$
$$= 97.3 \text{ g}$$

**Practice Exercises**

1. How many moles of $CaCl_2$ are required to make 0.500 L of a 0.500 mol/L solution?

2. What mass of AgCl is required to make 1.000 L of a 2.000 mol/L solution?

3. How many grams of $AgNO_3$ are required to make 2 L of a 0.100 mol/L solution? Would you put all of this into the flask? If not, explain why not.

3. Determining Concentrations of Solutions

If you are given the volume of a solution and this mass of solute you can calculate the concentration of the solution by using the same formula as given before.

$$C = \frac{n}{V}$$

It is important to first convert the mass into moles.

**Example**

Determine the concentration of a solution formed by dissolving 20.00 g or NaOH to make 0.200 L of solution.

molar mass of NaOH = 40.00 g/mol

$$n = \frac{20.00\ g}{40.00\ g/mol} = 0.5000\ mol$$

$$C = \frac{n}{V} = \frac{0.5000\ mol}{0.200\ L} = 2.50\ mol/L$$

4. Problem Solving Involving Concentrations

**Example**

An acid solution reacts with a base to produce a salt and water according to the reaction:

$$H_2SO_4 + 2NaOH \longrightarrow Na_2SO_4 + 2H_2O$$

If 0.250 L of a 0.200 mol/L solution of sulfuric acid ($H_2SO_4$) reacts with sufficient NaOH, what mass of water is produced?

Answer Step 1: Determine the number of moles of acid which reacted.

$$n = CV$$
$$= (0.200 \text{ mol/L})(0.250 \text{ L})$$
$$= 0.0500 \text{ mol}$$

Step 2: From the balanced equation 1 mol of $H_2SO_4$ yields 2 mol of $H_2O$ therefore 0.0500 mol of $H_2SO_4$ yields 0.100 mol of $H_2O$.

Step 3: The mass of 0.100 mol of $H_2O$ is

$$m = nM$$
$$= (0.100 \text{ mol})(18.02 \text{ g/mol})$$
$$= 1.80 \text{ g}$$

**Practice Exercises**

1. What would be the concentration of a $AgNO_3$ solution if 21.24 g of $AgNO_3$ was dissolved in water to make 0.750 L of solution?

2. A solution of $BaCl_2$ has a concentration of 0.0375 mol/L. What mass of $BaCl_2$ is dissolved in 2 L of solution?

5. Dilution of Solutions

Many solutions in a laboratory are in concentrated form and are diluted to required concentrations and volumes with distilled water. This saves much storage space and also delivery costs.

The important concept in dilution problems is that the number of moles of solute does not change when a solution is diluted. A simple example is if you have 1 L of a 1 mol/L salt solution you have 1 mol of salt dissolved. If you add water to make 10 L of solution you still have only 1 mol of salt (no salt was added) but now it is dissolved in 10 L of water.

A formula used in dilution problems is:

$C_iV_i = C_fV_f$ where $C_i$ = initial concentration

$C_f$ = final concentration

$V_i$ = initial volume

$V_f$ = final volume

**Examples**

1. 0.0500 L of a 0.800 mol/L $AgNO_3$ solution was diluted to make a volume of 1.000 L. Determine the final concentration.

Answer: $C_f = \frac{C_iV_i}{V_f}$

$$C_f = \frac{(0.800 \text{ mol/L})(0.0500 \text{ L})}{1.000 \text{ L}}$$

$$C_f = 0.0400 \text{ mol/L}$$

2. Determine the initial volume of a 3.000 mol/L $KMnO_4$ solution if it was diluted to produce 0.400 L of a 2.000 mol/L solution.

$$V_i = \frac{C_fV_f}{C_i} = \frac{(2.000 \text{ mol/L})(0.400 \text{ L})}{3.000 \text{ mol/L}}$$

$$V_i = 0.267 \text{ L}$$

**Practice Exercises**

1. Determine the concentration if 0.375 L of a 2.50 mol/L solution was diluted to 1.75 L.

2. Calculate the volume of a solution which resulted when 0.500 L of a 2.00 mol/L HCl solution is diluted to produce a 0.100 mol/L solution.

Exercises

1. If you were to prepare 200.0 mL of 0.375 mol/L AgCl solution, what mass of AgCl is required?

2. What mass of sodium iodide, NaI, is required to prepare 1.000 L of a 0.3 mol/L solution?

3. What mass of aluminum sulfate, $Al_2(SO_4)_3$ would be required to make $1.5 \times 10^3$ mL of a 0.05 mol/L solution?

4. Determine the concentration of a solution made by dissolving 34.00 g of $AgNO_3$ in water to make 0.400 L of solution.

5. 0.800 L of a $CuSO_4$ solution was found to contain 31.9 g of $CuSO_4$. What was the concentration of the solution?

6. Zinc reacts with hydrochloric acid to produce hydrogen gas and zinc chloride according to the equation.

$$Zn + 2HCl \longrightarrow ZnCl_2 + H_2$$

If 0.600 L of a 0.100 mol/L solution of hydrochloric acid reacts with sufficient zinc, determine:

(a) number of moles of zinc chloride produced.

(b) number of moles of hydrogen gas produced.

(c) the mass of zinc chloride produced.

7. Potassium chromate reacts with barium nitrate to yield a precipitate, barium chromate, and potassium nitrate according to the reaction

$$K_2CrO_4 + Ba(NO_3)_2 \longrightarrow BaCrO_4 + 2KNO_3$$

If 1.000 L of potassium chromate solution took part in the reaction to produce 126.70 g of barium chromate, calculate the concentration of the potassium chromate solution. (Hint: first determine the number of moles of $BaCrO_4$ produced)

8. What is the final concentration of a KI solution made by diluting 0.200 L of a 0.800 mol/L solution to 1.000 L?

9. To what volume must 15.00 mL of a 6.000 mol/L $H_2SO_4$ solution be diluted to produce a 0.250 mol/L solution?

10. What volume of a 2.000 mol/L solution of $Pb(NO_3)_2$ is required to produce 0.600 L of a 0.250 mol/L solution?

**Answers to Practice Exercises**

Page 15

1. 0.250 mol
2. 286.64 g
3. 33.98 g; No, if you are using a 1 L volumetric flask you will have to put 16.99 g into the flask, make the solution and repeat to make 2 L.

Page 16

1. 0.167 mol/L
2. 15.62 g (20 g correct to 1 significant figure)

Page 17

1. 0.536 mol/L
2. 10.0 L

**End of Lesson G**

# Chemistry Upgrading
# Lesson H

**DO NOT** send this lesson for correction

SOLUTIONS (Continued)

D. Solubility and Equilibrium

1. Unsaturated Solutions
2. Saturated Solutions
3. Supersaturated Solutions

E. Factors That Affect Solubility

1. Nature of Solute and Solvent
2. Temperature
3. Pressure

F. Experiment 2 - Bond Types and the Solution Process

## D. Solubility and Equilibrium

Gaseous solutions and many liquid solutions can dissolve into each other in all proportions, however, when a solid is added to a liquid, there is often a limit as to how much can dissolve.

It is evident that when one teaspoon of sugar is put in coffee, you can stir and dissolve the sugar completely. The same is true for two teaspoons of sugar. This would be an unsaturated solution. When you keep on adding sugar, you soon will come to the point where all of the sugar dissolves, but an additional grain of sugar would just fall to the bottom and would refuse to dissolve. At this point, the solution is saturated. Whenever you have a beaker of water, and a solid at the bottom which refuses to dissolve, you know you have a saturated solution of that particular solid. Naturally, there is a great difference in the amount of a specific substance you must add to get a saturated solution. Large amounts of table salt or sugar must be added to get a saturated solution, but small amounts of gases or different salts may be necessary to get a saturated solution of that particular substance.

In a saturated solution, there is an equilibrium between the undissolved solid at the bottom, and the dissolved particles in the solution. This is because as some dissolved solute particles crystallize, other undissolved particles go into solution. Since this occurs at equal rates, there is no net change in the amount of undissolved solute.

Supersaturated solutions contain greater amounts of solute than they could hold under normal conditions. For example, pop can hold much carbon dioxide under pressure, but less at atmospheric pressure. When the cap is removed, and you have atmospheric pressure on the pop, the pop is supersaturated with carbon dioxide, which is why the bubbles start escaping from the pop.

It is also possible to have a saturated solution of a substance with undissolved solids at the bottom, and then the remaining solids could be made to dissolve on heating. This saturated solution can then be cooled, and if no solute precipitates out, the solution would be supersaturated. Naturally, any supersaturated solution is unstable, and it doesn't take much to cause the excess solute particles to come out of solution. In the case of a supersaturated solid in solution, a single grain of the solute will trigger off the precipitation process.

## E. Factors That Affect Solubility

### 1. Nature of Solute and Solvent

Whether or not a solid dissolves in a liquid is a hard thing to predict, due to opposing tendencies of randomness and energy. Perhaps an analogy will make this clear. Suppose we have two hostage cases which have 100 hostages in each area. The conditions in both areas are exactly the same, and in both places all of the hostages want to escape equally badly. You might say that the tendency toward maximum randomness is the same in both cases. There are some differences however. One area has three 2 m high, barbed wire fences, 100 m apart, right around the area. In addition, there are several hungry and mean dogs between each of these fences. In the second case, there is only one 2 m high, barbed wire fence, and no dogs to contend with. Assume that the terrorists are the same in both cases. Although the willingness of the hostages to escape is the same in both cases, (tendency toward randomness), in the first case, it takes much more energy to escape than in the second case. Therefore a lot fewer hostages will escape from the first area as from the second area (due to energy requirements).

Exactly the same thing happens with molecules and solutions. In each case, the molecules want to be more random and free. However, instead of having barbed wire fences and dogs to contend with, the makeup of each molecule will determine whether or not a given molecule will dissolve in the water. For example, consider NaCl and CdS. In both cases, similar polar water molecules would exert forces trying to separate the Na from Cl, and the Cd from the S. The bonds between Cd and S are very much stronger than the bonds between Na and Cl. Therefore the NaCl will dissolve readily, but the CdS will not.

Randomness for a gas means different things than randomness for a solid. If a solid molecule wants to become more random, it would go into solution, where it is more free than if it stayed in the solid form. A gas molecule on the other hand, feels restricted by the solvent molecules, and would be more free in the gas phase. However, if it takes too much energy to achieve this random state, the molecules "decide" not to bother.

Whether or not a solute dissolves in a solvent also depends on the solvent. For example, water is highly polar, therefore it will dissolve highly polar or ionic substances like NaCl. Gasoline, however, which is not very polar, will not dissolve NaCl. Consider an alcohol with the formula $C_9H_{19}OH$. Due to the long hydrocarbon end, which is not very polar, this alcohol will not readily dissolve in the polar water molecules, but it will dissolve in relatively nonpolar gasoline molecules ($C_8H_{18}$). The conclusion that can be drawn from this is that like dissolves like. Polar solutions dissolve polar or ionic substances. Nonpolar solvents dissolve nonpolar substances.

As an analogy, pretend you have 100 copper pennies in a jar. By thoroughly shaking the jar, the 100 pennies will be completely mixed up. Now, add 30 magnetized steel pennies of the same mass as the copper into the jar. Now, when the jar is thoroughly shaken, all of the magnetized pennies will group together and will not "dissolve" in the copper pennies. The only thing that can take one magnetized steel penny away from the next steel penny is another magnetized steel penny. In this analogy, the copper molecules would be the nonpolar molecules, and the magnetized steel pennies would be the highly polar or ionic molecules. Carrying this analogy a bit further, you know that some ionic molecules dissolve in water better than other ionic molecules. Similarly, some steel pennies could be more highly magnetized than others. Also, thirty very weakly magnetized (not too polar) steel pennies would dissolve much more readily in the nonmagnetized (nonpolar) copper pennies. Thirty nonmagnetized zinc pennies of equal mass would of course be very evenly dispersed in the copper pennies, since both are not magnetized (nonpolar).

2. Temperature

According to Le Chatelier's Principle, if a stress is applied to a system, the system will readjust to reduce the effects of the stress. The stress of added heat will cause the molecules to move about more randomly. Generally, this means that gases are less soluble at higher temperatures, and solids are more soluble. However, this is not always the case.

It may seem odd that increasing the temperature of a solution may either increase or decrease the solubility of a given substance. All substances have more energy at higher temperatures, and hence intermolecular bonds and intramolecular bonds are more likely to break than at low temperatures. When something dissolves in water, two things occur:

(1) The bond between the substances is broken, and

(2) Water molecules are attracted to the individual solute particles.

Of course, a higher temperature weakens the bond between solute particles, but it would also weaken the "bond" between water molecules and the solute. If an increase in temperature causes (1) to weaken at a faster rate than (2), then solubility increases as the temperature increases. If (2) weakens at a faster rate at a higher temperature, then solubility decreases at a higher temperature.

3. Pressure

Since solids and liquids cannot be compressed, pressure has little effect here. However, pressure greatly affects gases. As a result, more gases will dissolve in a liquid if the pressure is higher. When the pressure is then removed, the gases come out of solution, as you have observed with pop.

Exercise 1

1. Define a saturated solution. _____

_____

_____

2. Does a saturated solution have to have undissolved solid at the bottom? _____ Explain. _____

_____

_____

_____

3. Define a supersaturated solution. _____

_____

_____

_____

4. (a) If the radio announcer says that the relative humidity is 100%, this would mean that the air is (unsaturated, saturated, or supersaturated) with water vapor at that temperature.

(b) An increase in temperature at this point would make the air _____

(c) A decrease in temperature at this point would make the air _____, unless rain would start falling immediately.

5. When a bottle of pop is opened, the $CO_2$ would fizz out until

   (a) absolutely no $CO_2$ is left
   (b) the solution is saturated with $CO_2$ _____

6. When a cup of water is heated, bubbles start to form even before the water starts to boil. Explain in terms of solutions.

   _____

   _____

7. Oil based paints cannot be washed off with water, but varsol will take the paint off. From this information, what do you conclude about the relative polarity of the various molecules? _____

   _____

   _____

   _____

8. Raising the temperature always (increases or decreases) the tendency towards randomness, and (strengthens or weakens) the bonds between the various molecules.

9. Solids are (more, less) random in the dissolved state than in the solid state.

10. Gases are (more, less) random in the dissolved state than in the gaseous state.

11. The carbon tetrachloride molecule ($CCl_4$) is nonpolar since it is symmetrical in shape. Would you expect table salt to dissolve in $CCl_4$ ? _____

    Would you expect water to dissolve in $CCl_4$ ? _____

    Would you expect butane ($C_4H_{10}$) to dissolve in $CCl_4$ ? _____

12. You have a cup which contains a saturated solution of NaCl in water, with 1 cm of NaCl lying at the bottom. Increasing the pressure would

    (a) make more salt dissolve in the water.
    (b) make more salt precipitate out of the solution.
    (c) have virtually no effect on the amount of salt in solution. _____

13. In the bottle at the left, what would happen to the number of molecules of $CO_2$ dissolved in the liquid if you were to inject $10^5$ molecules of $CO_2$ into the top, and leave everything else unchanged?

(a) more $CO_2$ would dissolve in the liquid
(b) some $CO_2$ would come out of the liquid
(c) there would be virtually no change in the amount of $CO_2$ in the liquid _____

14. A small lake contains enough dissolved oxygen to sustain trout and other fish, even at the warmest time of the year. What would happen to the temperature of the water if a nuclear power plant were to use the water of this lake for cooling purposes?

_____

What would happen to the amount of oxygen the water could then dissolve?

_____

What would then happen to the trout population? _____

15. A closed jar half filled with water contains dissolved oxygen. If both the pressure and temperature are increased, how will the amount of dissolved oxygen be affected?

(a) increase
(b) decrease
(c) not enough information is given. _____

16. Fifty grams of a solid are put into some water which can only dissolve thirty grams of this substance. After a while, thirty grams are in solution, and twenty grams are at the bottom. After an hour, the amounts in each phase would be the same, but would the same molecules still be in each phase as an hour ago? _____ Explain _____

_____

_____

_____

Let us assume that 15 g of sodium thiosulfate can dissolve in 3 $cm^3$ of water at 75°C. Furthermore, assume that if the temperature is increased, more $Na_2S_2O_3 \cdot 5H_2O$ can dissolve in water. With this information, answer the questions below. Your volume of water is always 3 $cm^3$.

Exercise 2

1. If we assume that 15 g of sodium thiosulfate can dissolve in 3 $cm^3$ of water at 75°C, what kind of a solution would you have if 15 g are dissolved in 3 $cm^3$ of water at 90°C?

   (a) unsaturated
   (b) saturated
   (c) supersaturated _____

2. If 15 g were dissolved at 75°C, your solution would be

   (a) unsaturated
   (b) saturated
   (c) supersaturated _____

3. If 15 g were dissolved at 23°C, the solution would be

   (a) unsaturated
   (b) saturated
   (c) supersaturated _____

4. If 15 g of $Na_2S_2O_3 \cdot 5H_2O$ were heated to 50°C, and you had some of the substance at the bottom of the test tube in the undissolved state, the water above would have a(n)

   (a) unsaturated solution of $Na_2S_2O_3 \cdot 5H_2O$
   (b) saturated solution
   (c) supersaturated solution _____

5. Which solution is most unstable?

   (a) unsaturated
   (b) saturated
   (c) supersaturated _____

6. If sodium thiosulfate cools the water initially when it dissolves, how would you expect the temperature of the water to change when the sodium thiosulfate starts to recrystallize after the solution has been supersaturated? _____ Would these temperature changes indicate that energy is conserved? _____

7. Since sodium thiosulfate dissolves so well in water, the sodium thiosulfate molecules are (polar or nonpolar).

8. Sodium thiosulfate would (dissolve or remain undissolved) if placed into cooking oil.

9. Would raising the pressure have a significant effect on the amount of $Na_2S_2O_3 \cdot 5H_2O$ that could be dissolved? _____ Why? _____

_____

_____

_____

10. The initial dissolving process on the addition of heat could be represented by the equation:

$$H_2O + Na_2S_2O_3 \cdot 5H_2O_{(s)} + \text{heat} \longrightarrow \text{dissolved } Na_2S_2O_3 \cdot 5H_2O.$$

Rewrite the equation that would represent what happens after the single crystal is added to the supersaturated solution?

$$Na_2S_2O_3 \cdot 5H_2O_{(aq)} \longrightarrow \_\_\_\_\_ \quad \_\_\_\_\_ \quad \_\_\_\_\_$$

By the way, if there were no dust particles or other "seeds" in the air, the air could easily become supersaturated with water vapour since the water molecules need something to start building on. You may have heard of efforts to induce rain to fall in certain areas. This is done by injecting some seed particles such as AgI into the air onto which the water vapour in the air can start forming droplets.

11. The $5H_2O$ portion in $Na_2S_2O_3 \cdot 5H_2O$ is called the water of hydration. Would such a large amount of sodium thiosulfate have dissolved in water if it didn't contain water of hydration in the formula? _____

12. Of the 15 g $Na_2S_2O_3 \cdot 5H_2O$, how many grams were water? _____

F. Experiment 2 - Bond Types and the Solution Process

Things that dissolve in (polar) water will not necessarily dissolve in (nonpolar) oil. In this experiment, you will see how sugar, salt, oil, and water dissolves in water, and how these same substances dissolve in oil.

Below, are the headings for the table you will be asked to fill in. Write "yes" if the substance on the left dissolves in the substance at the top, and write "no" if the two substances don't dissolve in each other. In all cases, shake or stir vigourously to give the solute a good chance to dissolve in the solvent. Don't use too much solute because if an excessive amount is used, you are bound to get a solution where some solute remains at the bottom, and it may be hard to tell if any of it has dissolved. It is much easier to tell what is happening if you add 4 grains of salt in water and have them dissolve, rather than adding a lot and not being sure whether or not some did dissolve. Also, small amounts of solute don't require as much shaking or stirring.

Follow the procedure in such a way that after you are finished with one solvent and solute, you clean the glass thoroughly, and work your way down with the next solvent.

Data Table:

| | Solvents | |
|---|---|---|
| Solutes ↓ | water | oil |
| sugar | | |
| oil | | |
| salt | | |
| water | | |

Exercise 3

1. Water is (polar, nonpolar), therefore it should dissolve (polar, nonpolar) substances.

2. (a) On the basis of their solubilities in water, which substances are polar?

   _____

   (b) Which are nonpolar? _____

3. Oil is (polar, nonpolar), therefore it will dissolve (polar, nonpolar) substances.

4. Would you expect iodine to dissolve in gasoline? _____ Why? _____

   _____

5. Would you expect KCl to dissolve in gasoline? _____ Why? _____

   _____

6. Would you expect gasoline to dissolve in cooking oil? _____

   Why? _____

   _____

END OF LESSON H

# Chemistry Upgrading
# Lesson I

**DO NOT** send this lesson for correction

A. Introduction
B. Alkanes
C. Alkenes
D. Alkynes
E. Physical and Chemical Properties of Aliphatics
F. Cyclic Compounds
G. Aromatics
H. Alcohols
I. Organic Acids
J. Esters
K. Additional Information

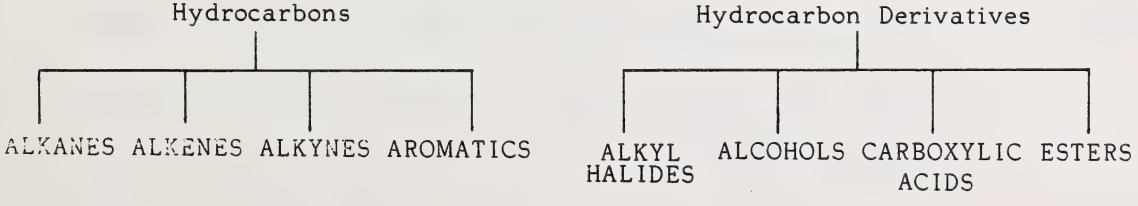
## A. Introduction

The most important element on and around the planet earth is carbon. The study of carbon and its compounds is so important in our daily lives that we treat it under a separate branch of chemistry called organic chemistry. Early in the history of chemistry, organic chemistry was thought of as the chemistry of substances from living things, both plant and animal. It was found later, that almost all of these substances were built from chains and rings of carbon atoms. Today, many compounds of carbon are man-made or synthesized. In 1828, a German chemist Freidrick Wohler, syntheized the first organic compound called urea. His discovery led the way to the synthesis of many other useful compounds such as plastics, drugs, fibres, food-products, detergents and perfumes. Manufacture of these compounds has given rise to the growth of many new industries which require millions of tonnes of raw materials each year. Can you imagine what our world would be like without the element carbon? It is little wonder that scientists are continually searching for traces of carbon on other planets.

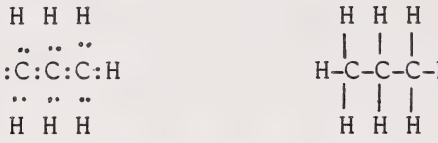
Organic fuels include gasoline, wood, oil, and natural gas. Organic drugs include vitamins, aspirin, penicillin, and tranquilizers. Organic fibers include wool, cotton, linen and nylon. All of the above substances contain carbon and hydrogen and often oxygen and nitrogen and other substances may also be present.

Fossil fuels are a major source of energy in our society. As these supplies are dwindling, new sources of energy must be made practical for future generations.

Below is a diagram representing the types of hydrocarbons we will study in this lesson.



There are several ways that a compound can be represented. You can write the molecular formula, or the structural formula, or the electron dot formula. For the electron dot formula, keep in mind that each carbon wants to be surrounded by an octet of eight electrons and that each hydrogen atom wants two electrons. When writing the structural formula, each dash represents a bond of two electrons. Therefore each carbon should have 4 dashes and each hydrogen should have one dash. Below, is an illustration of how propane can be shown.



$CH_3CH_2CH_3$ $C_3H_8$

The naming of hydrocarbons has been greatly simplified through the development of a systematic method of nomenclature by the International Union of Pure and Applied Chemistry (IUPAC). We will focus on the first ten carbon chain hydrocarbons. Since the Greek prefixes are used throughout it is advisable that the student memorize the first ten prefixes below. This will make this lesson much easier to master.

| Prefix | Meaning |
|---|---|
| meth | 1 |
| eth | 2 |
| pro | 3 |
| but | 4 |
| pent | 5 |
| hex | 6 |
| hept | 7 |
| oct | 8 |
| non | 9 |
| dec | 10 |

## B. Alkanes (Saturated Organic Compounds)

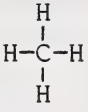
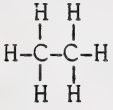
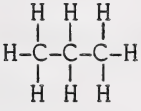
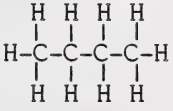
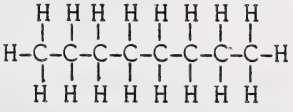
This group of compounds, straight chain and branched chain, has the general formula:

$$C_nH_{2n+2}$$

Each carbon atom has four single bonds, one bond to each of four different atoms. The rules for naming straight chain alkanes are simple:

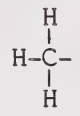
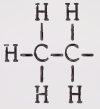
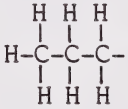
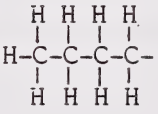
1. the prefix indicates the number of carbon atoms in the compound
2. the prefix is followed by **ane** which indicates a saturated hydrocarbon

Below is a table of some of the straight chain alkanes.

| Name | Structured Formula | Molecular formula |
|---|---|---|
| methane |  | $CH_4$ |
| ethane |  | $C_2H_6$ |
| propane |  | $C_3H_8$ |
| butane |  | $C_4H_{10}$ |
| octane |  | $C_8H_{18}$ |

Alkyl groups have the general formula: $C_n H_{2n+1}$

Below is a table of some alkyl groups (notice the difference between the straight chain alkanes).

| Name | Structural Formula | Molecular Formula |
|---|---|---|
| methyl |  | $CH_3$ |
| ethyl |  | $C_2H_5$ |
| propyl |  | $C_3H_7$ |
| butyl |  | $C_4H_9$ |

Branched chain alkanes are named in the following manner:

1. determine the number of carbon atoms in the longest chain of carbon atoms.

2. number of carbon atoms starting at the end closest to the branch.

3. give the number of the carbon to which the branch is attached, give the name of the alkyl group followed by the name of the longest carbon chain.

4. If more than one alkyl group is present:

   (i) if the alkyl groups are the same use the Greek prefixes

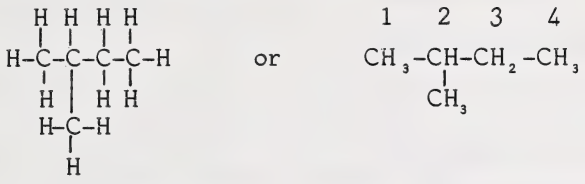
   mono - 1
   di - 2
   tri - 3
   tetra - 4
   penta - 5

   to indicate the number of that group

   (ii) if the alkyl groups are different, order them according to complexity or alphabetical order.

**Examples**

(a) Provide the name for the hydrocarbon below.

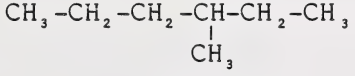
1. 

2. (Notice we have numbered the longest carbon chain.)

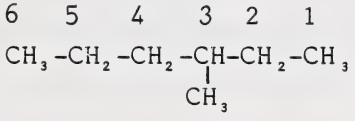
3. 2-methylbutane

the methyl group is attached to carbon #2 of the straight chain (butane).

(b) Provide the name for the hydrocarbon shown below:

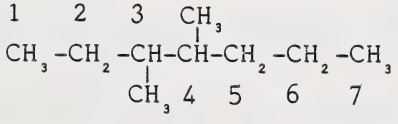


1,2) number the longest chain, (starting at the end closest to the branch.



3. 3-methylhexane

(c) Provide the name for the hydrocarbon.



1,2) done above

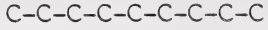
3,4i)

3,4 - dimethylheptane

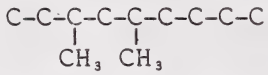
(d) draw the structural formula of the hydrocarbon

3,5 dimethyl - 4 ethylnonane

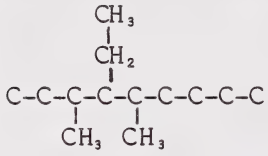
1. draw nonane carbon chain



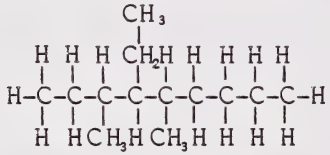
2. add methyl to carbon 3,5



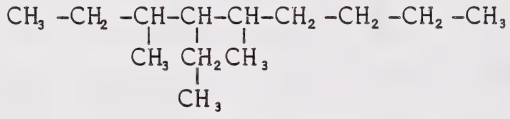
3. add ethyl to carbon 4



4. fill in the remaining possible carbon bonds with hydrogen.

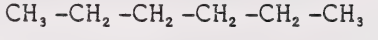


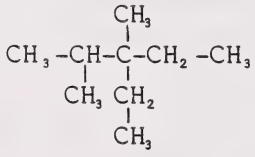
You can write this as

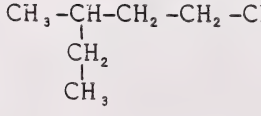


## Practice Exercises

1. Provide the name of the hydrocarbons:

(a) 

(b) 

(c) 

remember the longest chain

2. Draw the structural formula for each:

   (a) 3,5 - dimethyl - 4 - ethyloctane

   (b) 3 ethylheptane

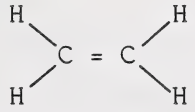
   (c) pentane

## C. Alkenes (Unsaturated Organic Compounds)

Straight chain and branched alkanes all consisted of single bonds between atoms. There exist many compounds which contain double bonds (two pairs of electrons are shared between two carbon atoms). These organic compounds are called alkenes. The general formula for alkenes is

$$C_nH_{2n}$$

Notice alkenes have two less hydrogen atoms than the alkane with the same number of carbon atoms. The simplest alkene is ethene $C_2H_4$ whose structure is shown below.



Notice each carbon atoms has four pairs of electrons (four bonds).

Nomenclature of the alkenes follows these rules:

1. ending **ene** is used to indicate the double bond.

2. Identify the longest chain containing the double bond. Number the carbon atoms in the chain with the lowest numbers applying to the double bond. (Start at the end closest to the double bond.)

3. name the branches as with alkanes; name the stem with a number indicating the position of the double bond.

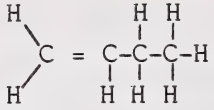
### Examples

1. $CH_2 = CH_2$

ethene (also common name is ethylene)

| eth means | ene indicates |
|---|---|
| two carbon | a double |
| atoms | bond |

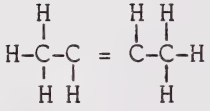
2.



1 - butene

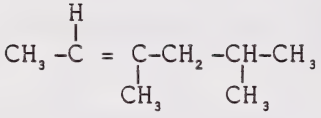
the 1 indicates the double bond occurs in the first carbon atom.

3.



2 - butene

4.



3,5 - dimethyl - 2 - hexene

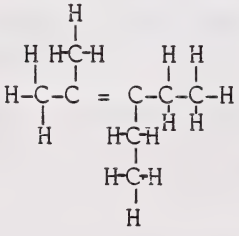
## Practice Exercises

3. Provide the names of the following compounds

(a) $CH_2 = CH-CH_3$

(b) $CH_3-CH_2-CH = CH-CH(CH_2CH_3)-CH_2-CH_3$

(c)



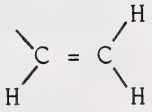
4. Draw the structural formula for each:

(a) 2,2 - dimethyl - 3 hexene

(b) 2 - methyl - 3 - ethyl - 2 octene

(c) 2 - methyl - 1 - butene

Interest the group



is called a vinyl group and is used in the production of a tremendous variety of plastics.

## D. Alkynes (Unsaturated Organic Compounds)

Compounds which contain triple bonds between two carbon atoms are called alkynes. The general formula for alkynes is

$$C_nH_{2n-2}$$

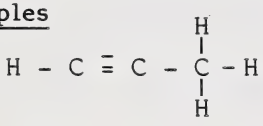
The simplest alkyne is ethyne shown below

$$H - C \equiv C - H$$

### Nomenclature of Alkynes

1. ending yne is used to indicate the triple bond
2. determine the longest chain containing the triple bond and number it with the lowest numbers assigned to the carbon with the triple bond.
3. name the branches as in alkenes.

### Examples

1. 

   propyne

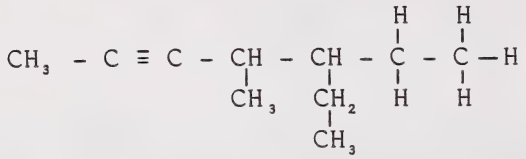
2. $CH \equiv C - CH_2 - CH_3$

   1 - butyne

3. $CH_3 - C \equiv C - CH_3$

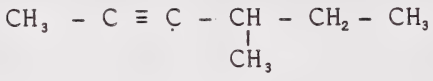
   2-butyne

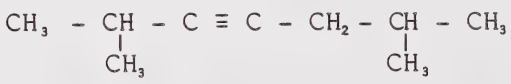
4. 4 - methyl - 5 - ethyl - 2 - heptyne



**Practice Exercises**

5. Provide the names of the compounds:

a) 

b) 

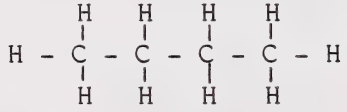
6. Draw the formula of each

a) 4, 4 - diethyl - 1 hexyne

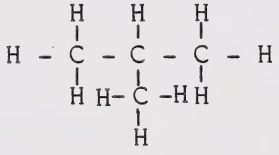
b) 3 - octyne

**Isomerism in Alkanes, Alkenes, and Alkynes**

Compounds which have the same molecular formula but different structures are called isomers. These isomers have different physical and chemical properties and hence different uses. The simplest example is found in $C_4H_{10}$ . Look at the two structures below and verify for yourself they are both $C_4H_{10}$ .



butane



**methylpropane**

**Practice Exercise**

7. Draw the structures and provide the names for five noncyclic isomers of $C_5H_{10}$ .

## Physical and Chemical Properties of Aliphatics

The physical properties of alkanes, alkenes, and alkynes are in general

- gases when the compound contains 4 or fewer carbon atoms e.g. methane (natural gas)
- liquids between 5 and 16 carbon atoms e.g. octane (gasoline)
- solids with more than 16 carbon atoms e.g. paraffin ($C_{25}H_{52}$)

Chemical properties which are common to all three are

- relatively unreactive under ordinary conditions
- react with oxygen at high temperatures to produce $CO_2$ and $H_2O$ (when combustion is complete).

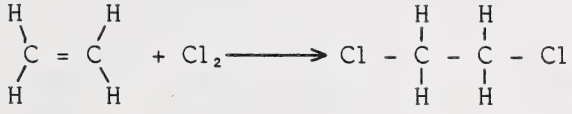
Dissimilar chemical properties:

- Alkanes undergo substitution reactions where a hydrogen atom is replaced by another atom or a group of atoms.

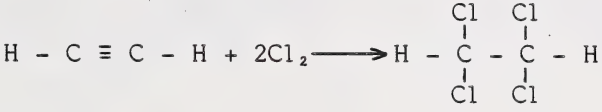
eg)

- Alkenes and alkynes undergo addition reactions where the multiple bond is broken and other atoms are bonded to the carbon atoms.

eg)



eg)



- The reactivity of the organic compounds can be compared where the reactivity of

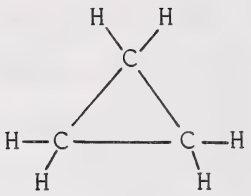
**alkanes < alkenes < alkynes**

since the multiple bonds between carbon atoms are not as stable as single bonds.

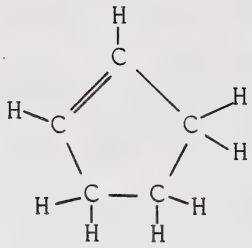
## F. Cyclic Compounds

Another group of hydrocarbons have a closed chain or ring structure. The prefix **cyclo** indicates a ring compound. Ring compounds can be either saturated or unsaturated.

Examples:



cyclopropane

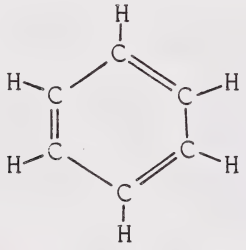


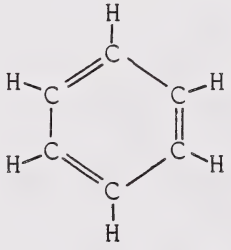
cyclopentene

The chemical and physical properties of the cyclic compounds are similar to their corresponding alkane or alkene except that the cyclic compounds are more reactive due to the strained nature of the carbon ring.

## G. Aromatics

There is another important group of cyclic compounds that is different from the group previously described. The simplest of this group is the compound benzene, $C_6H_6$. This cyclic compound has six carbon atoms in a ring with a single hydrogen atom attached to each carbon atom. The structural formula shows three double bonds, which may be represented by the following diagrams.





Each structure shows three of the carbon-carbon bonds are double bonds and three are shown as single bonds. But experimental evidence indicates this is not true. Any one of the six carbon-carbon bonds in benzene is the same as any other. Apparently each carbon atom shares electrons equally with each adjacent carbon atom. This makes it difficult to represent bonding for benzene by our usual line drawing. We will use the common symbol for benzene as shown below. The circle represents the evenly distributed electrons among the six carbon atoms.

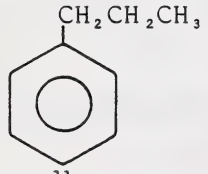
One must remember that the carbon-carbon bonds are actually the same between all carbon atoms and that these bonds exhibit properties unlike those of just single or just double bonds found in other compounds.

The compound benzene and its derivatives are called **aromatic** compounds, because of the fragrance of many of their derivatives. Some are pleasantly fragrant while others are not. Many are hazardous to health and should be handled with extreme care. Benzene is the starting material in the manufacture of many organic dyes, drugs and explosives.

Benzene undergoes substitution reactions not addition reactions. One or more of the hydrogen atoms is replaced by alkyl groups or a variety of other atoms or molecules. The benzene molecule is quite stable and remains intact.

Nomenclature:

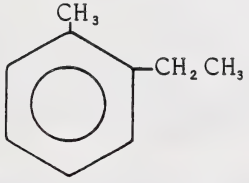
Type 1: a single alkyl group attached by an end carbon to the benzene ring.



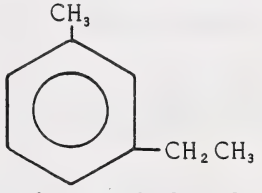

propylbenzene

Name the alkyl group and the benzene ring.

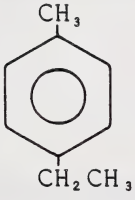
Type 2: two alkyl groups each attached by an end carbon




1 - methyl - 2 - ethylbenzene
or 1 - ethyl - 2 - methylbenzene
or o - methylethylbenzene

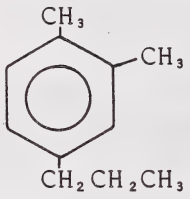



1 - methyl - 3 - ethylbenzene
or 1 - ethyl - 3 - methybenzene
or m - methylethylbenzene




1 - methyl - 4 - ethylbenzene
or 1 - ethyl - 4 - methylbenzene
or p - methylethylbenzene

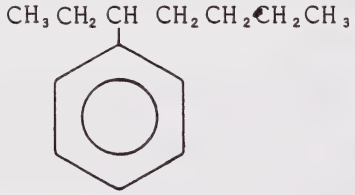
Type 3: three alkyl groups each attached by an end carbon.



1- 2 - dimethyl - 4 propylbenzene

In types 2 and 3 the carbon atoms in the benzene ring are numbered so as to provide the lowest numbers possible for the alkyl group which are attached to the ring.

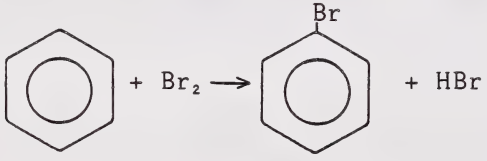
Type 4: benzene attached to an alkyl group at other than an end carbon



3 phenylheptane

Number the carbon atoms in the alkyl group so as to provide the lowest number, give the number of the carbon to which the benzene ring is attached, prefix the name of the alkyl group with phenyl.

Benzene undergoes substitution reactions with compounds other than hydrocarbons. An example is the reaction below:



bromobenzene

An excellent source of benzene is coal (actually coal tar) hence Alberta has tremendous potential for production of the great variety of materials in demand today such as medicines, insecticides, perfumes, and plastics to name only a few.

## H. Alcohols

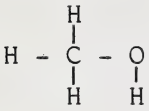
An alcohol is produced when a hydrocarbon is reacted with the hydroxide ion. The bond between the carbon atom and the oxygen atom is covalent. The alcohol molecule is nonionic and quite stable. Because the O-H bond is polar and the rest of the molecule is nonpolar alcohols exibit an unusual property; they can dissolve both ionic and molecular solutes. The longer the carbon chain, the less polar the alcohol.

Nomenclature:

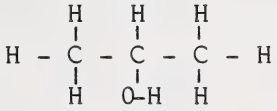
1. name the hydrocarbon; drop the e from the hydrocarbon name and add ol. (diol, triol indicate 2 and 3 hydroxyl (OH) groups respectively).

2. number the carbon atoms so as to arrive at the lowest possible number for the hydroxyl group.

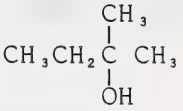
Examples:



methanol



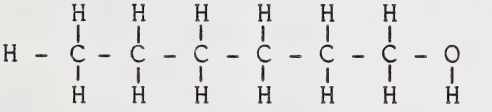
2 - propanol



2 - methyl - 2 - butanol

To simplify writing structural formulas of alcohols, acids, and esters, the alkyl group is often written simply as $C_nH_m$ or R.

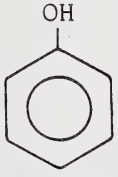
Examples:



or $C_6H_{13}OH$ or ROH

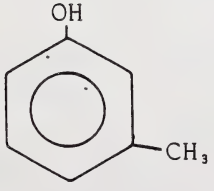
1 octanol can be written ROH where R is understood to be $C_8H_{17}$ in this particular case.

When benzene is bonded to the OH group the name given to this group is phenol.



phenol

Example



3 methylphenol

There is another system used which labels the OH group as a hydroxy. In the example above, 3 methylphenol could be called:
1 - methyl - 3 - hydroxybenzene.

Practice Exercises

8. Draw the structure of each

a) 1, 2 - ethanediol
b) 2 - pentanol
c) 3 - methyl - 2 - hexanol
d) 1 - methyl - 3 - propylbenzene
e) 2 - phenylpentane
f) cyclohexanol

9. Provide the name for each:

a) $CH_3CH_2\underset{\displaystyle OH}{\underset{|}{CH}}\ CH_3$

_____

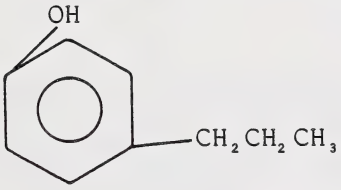
b) $C_3H_7OH$

_____

c) $CH_3\underset{\displaystyle CH_3}{\underset{|}{CH}} - CH_2 - CH_2 - CH_2 - CH_2OH$

_____

d)



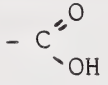

_____

e) just for fun

_____

## I. Organic Acids

Organic acids (also called carboxylic acids) have a carboxyl group in the molecule as shown below



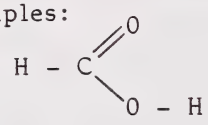

and is commonly written COOH. The physical properties of carboxylic acids are as follows:

- higher melting points compared to hydrocarbons and alcohols with an equal number of carbon atoms (this is due to hydrogen bonding between molecules).
- short carbon chain acids (4 or less carbons) are very soluble in water. Longer chain acids are not very soluble in water.
- they are weak acids and have properties similar to inorganic acids.

Organic acids are often found in fruits and living organisms. Because they are so common and known for so long, most have trivial (nonsystematic) names. IUPAC rules govern nomenclature of the carboxylic acids but the trivial names may also be used.

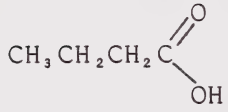
Nomenclature: drop the e from the parent hydrocarbon, add oic (dioic, trioic if two or three carboxyl groups are present) and add the word acid.

Examples:

1. 

methanoic acid (trivial name is formic acid)

from methane (one carbon).

2.  or $C_3H_7COOH$

butanoic acid (trivial name butyric acid)

Carboxylic acids are produced by oxidizing alcohols. An example of this is the production of acetic acid (ethanoic acid). This acid is the acid found in vinegar. Carboxylic acids are responsible for the tart (and often pleasant) taste of many fruits such as grapefruit, orange and lemons. Carboxylic acids behave chemically as do inorganic acids.

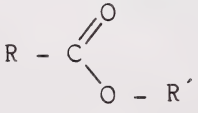
Example: ethanoic acid reacts with zinc.

$$ZN + 2CH_3COOH \longrightarrow Zn(CH_3COO)_2 + H_2$$

## J. Esters

Esters are prepared by reacting an organic acid and an alcohol. Some esters have pleasant tastes and odors, others do not. Esters are used in artificial flavors and scents such as peach, orange, strawberry, raspberry, etc. Virtually any flavour or scent can be produced. Animal fats and vegetable oils are complex esters.

Esters have the general formula

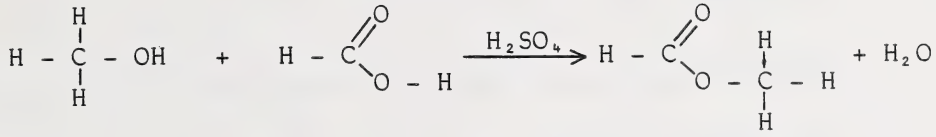


where R and R´ represent alkyl groups to simplify writing out the carbon chains.

Nomenclature:

- reduce the alcohol name to an alkyl name and give this first
  eg. methanol becomes methyl
- drop the ic from the acid name, add ate and give this name second.

Example: methanol reacts with methanoic acid to produce methyl methanoate and water.

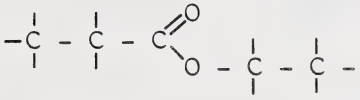



methanol methanoic acid methyl methanoate

**Practice Exercises**

10. Provide the systematic name for each:

a) $CH_3CH_2CH_2CH_2COOCH_3$

b) 

11. Give the structural formula for each:

a) ethyl butanoate

b) hexanoic acid

12. Write out the structural formulas of reactants and products when sodium hydroxide reacts with butanoic acid to produce sodium butanoate and water.

**Exercises**

1. Which element determines whether or not a substance is classified as organic?

_____

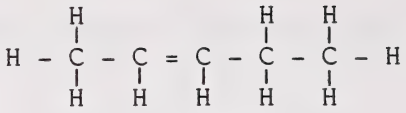
2. For each group, provide the names of three examples of the following groups.

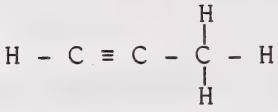
| Organic Fuels | Organic Drugs | Organic Fibers |
|---|---|---|
| _____ | _____ | _____ |
| _____ | _____ | _____ |
| _____ | _____ | _____ |

3. Give the systematic name for each of the following organic structures:

a) $C_6H_{14}$

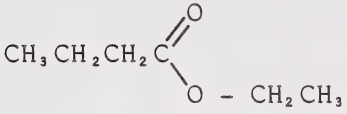
_____

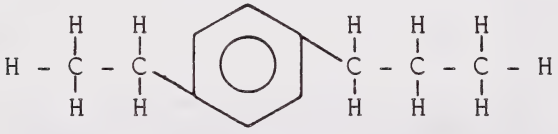
b) 

c) 

d) $C_3H_7OH$

e) $C_4H_9COOH$

f) 

g) 

4. Draw the structural formula for each.

a) propyl butanoate

b) pentanoic acid

c) cyclohexane

d) 2 - butanol

e) 2 - methyl - 3 - ethyl - octane

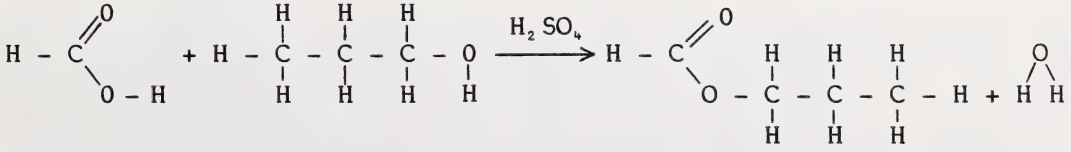
f) 3 - methyl - 2 - heptene

g) propyne

h) 1, 3 - dibromobenzene

5. For each reaction, provide the complete structural equations and the names of reactants and products..

eg.) methanoic acid + 1 - propanol $\xrightarrow{H_2SO_4}$ **propyl methanoate + water**



a) ethanoic acid + magnesium ⟶ _____ + _____

b) hexanoic acid and hexanol $\longrightarrow$ _____ + _____

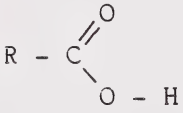
c) 2 - heptene and bromine ($Br_2$) $\longrightarrow$ _____

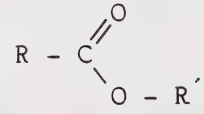
## K. Additional Information

There are four reasons why there are such a large number of carbon compounds. First of all, you can have chains of various lengths. Secondly, carbon atoms can share one, two, or three pairs of electrons. Thirdly, different isomers are possible. And lastly, different functional groups can be substituted for hydrogen.

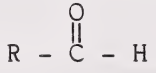
Summary of functional groups

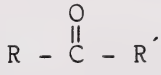
Alcohol - R - O - H

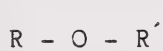
Acid - 

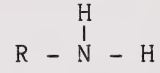
Ester - 

You are not required to know the following information, but you might want to read it for interest sake. We studied three possible functional groups which could be added to a hydrocarbon. As a result, we saw the structure for alcohols, organic acids, and esters. Additional compounds that you may come across are aldehydes, ketones, ethers, and amines. The structural formulas for these compounds are as follows:

aldehyde 

ketone 

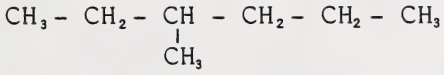
ether 

amine 

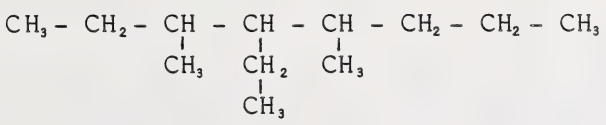
## Answers to Practice Exercises

1. a) hexane
   b) 2, 3 - dimethyl - 3 - ethylpentane
   c) 3 methylhexane

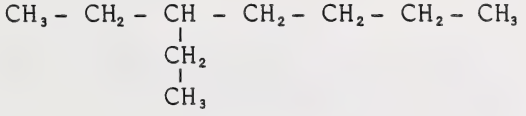
   Note: the answer is not 2 ethylpentane since the molecule can be drawn as



2. a)



   b)



   c) $CH_3 - CH_2 - CH_2 - CH_2 - CH_3$

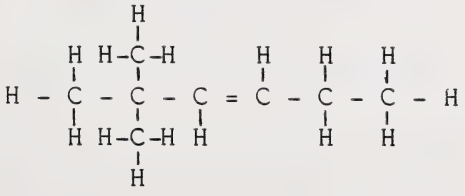
3. a) propene

   Note: 1 propene is not the answer since $CH_3 - CH = CH_2$ is exactly the same molecule just turned 180°. If you number the chain, the lowest number is on the right and it would still be propene therefore the number one is not required.

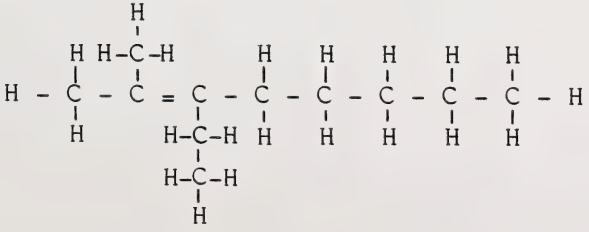
   b) 5 - ethyl - 3 - heptene

   c) 2 - methyl - 3 - ethyl - 2 - pentene

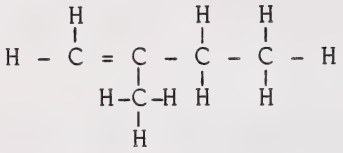
4. a)



   b)



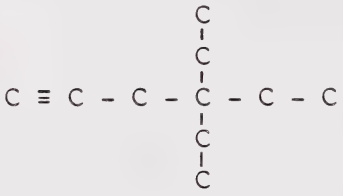
c)



5. a) 4 - methyl - 2 - hexyne

b) 2, 6 - dimethyl - 3 - heptyne

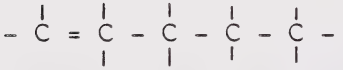
6. a)



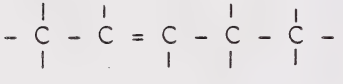
b) C - C - C ≡ C - C - C - C - C

Note: the hydrogen atoms and the bonds are omitted in the structural formula to save space only.

7. a)



b)



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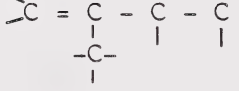

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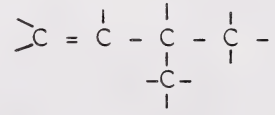
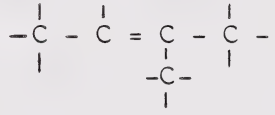
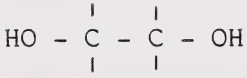


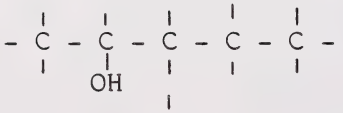
e)



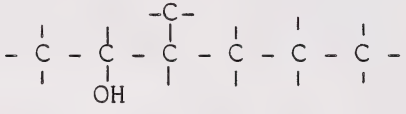
8. a)



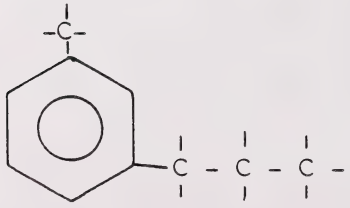
b)

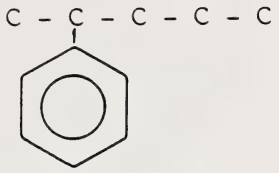


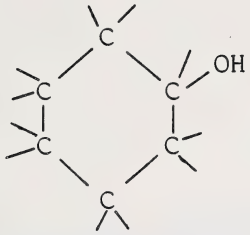
c)



d)

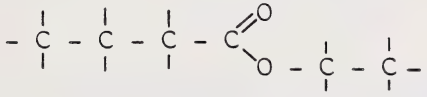


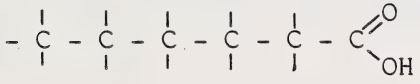
e) 

f) 

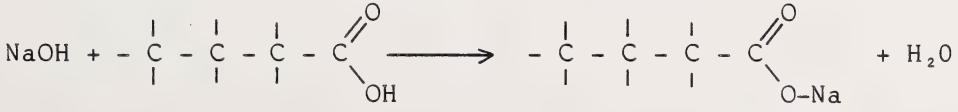

9. a) 2 - butanol
   b) propanol
   c) 5 - methyl - 1 - hexanol
   d) 4 - propylphenol
   e) 1, 2 - hydroxychickenwire

10. a) methyl pentanoate
    b) ethyl propanoate

11. a) 

    b) 

12. sodium hydroxide and butanoic acid ⟶ sodium butanoate and water



# Chemistry Upgrading
# Lesson J

**DO NOT** send this lesson for correction

CONTENTS

A. Introduction to Laboratory Work
1. Preparing for Laboratory Work
2. Apparatus
3. Laboratory Safety

B. Laboratory Techniques and Procedures
C. Experiment 3 - Qualitative and Quantitative Observations
D. Graphs
E. Experiment 4 - Melting and Freezing of Solids
F. Tests for Some Common Gases.

## A. Introduction to Laboratory Work.

In any science course, laboratory work is important. However, due to the nature of correspondence education, it is not possible to do all experiments since not all equipment can be sent by mail. Also, procedures may have to be modified somewhat to suit a correspondence course. In this lesson, you will learn a few basic things about various aspects in laboratory work.

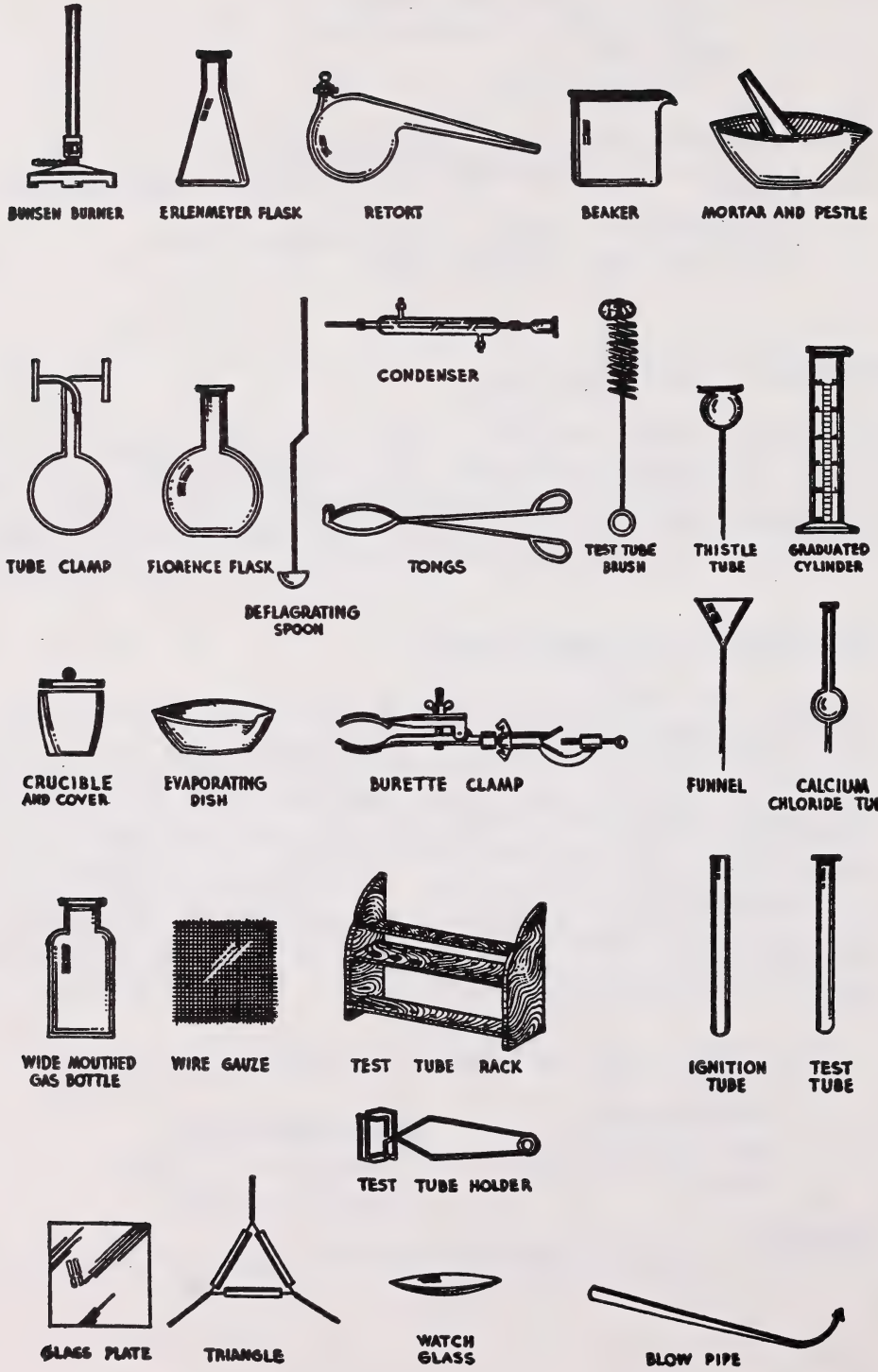
### 1. Preparing for Laboratory Work

It is very important that you understand the directions and purpose of each experiment before starting to work on it. It is a good idea to read the procedure through once before starting the experiment. As you read, decide what data you will have to record and prepare data tables. Then when you start the experiment, your observations can be recorded systematically, which makes the record meaningful for future use. Your record of experiments should be a complete but brief outline of the experiment. It is not necessary to copy out the instructions. What you did and what you saw are noted under the heading: "Record". The more precise your work is, the more valuable it will be.

In recording, use the passive voice and be impersonal.

| Recommended | Not Recommended |
|---|---|
| The test tube was heated. | I heated the test tube.<br>Heat the test tube. |

### 2. Apparatus - see list of common laboratory apparatus enclosed on page 2 of this lesson.


BUNSEN BURNER
ERLENMEYER FLASK
RETORT
BEAKER
MORTAR AND PESTLE
TUBE CLAMP
FLORENCE FLASK
CONDENSER
TONGS
DEFLAGRATING SPOON
TEST TUBE BRUSH
THISTLE TUBE
GRADUATED CYLINDER
CRUCIBLE AND COVER
EVAPORATING DISH
BURETTE CLAMP
FUNNEL
CALCIUM CHLORIDE TUBE
WIDE MOUTHED GAS BOTTLE
WIRE GAUZE
TEST TUBE RACK
IGNITION TUBE
TEST TUBE
TEST TUBE HOLDER
GLASS PLATE
TRIANGLE
WATCH GLASS
BLOW PIPE

3. Laboratory Safety

Following, are some things to keep in mind when you are doing experiments, since certain precautions must always be taken in laboratory work to avoid accidents or damage to clothes or equipment.

1. Follow the directions carefully.

2. Do not combine chemicals except according to the directions in the experiment. Have a large clear table to work on.

3. Keep your working area clean and organized.

4. Wear a protective laboratory coat or an old shirt over your clothes.

5. When you have opened a container and removed the amount of a reagent needed for the reaction, close the container and move it back on the table, where it is out of your way and you cannot accidentally brush it onto the floor.

6. Do not taste any of the chemicals. Do not leave them where they can be reached by children or pets.

7. If you spill any reagent on your skin or clothing, rinse immediately with plenty of cool water. If you spill a reagent on the table, wipe it up immediately and rinse the surface with a wet cloth.

B. Laboratory Techniques and Procedures

There are many laboratory techniques and procedures which you would have to be familiar with if you were to study this course in a classroom. However, there is little sense in going through things which you may either never need to know, or which you may need when you take further chemistry courses. Should you later need to know things like how to operate a bunsen burner, you can always learn it then. In this course, we will give a limited discussion of laboratory techniques and procedures.

There are many things that you can use to measure the volume of a liquid. Whichever apparatus you use, always be sure that you read the mark with your eye directly opposite the top of the liquid, and not above or below the liquid level. Also, if you have a meniscus, read the liquid level from the flat portion, and ignore the edges which may go either up or down at the edges.

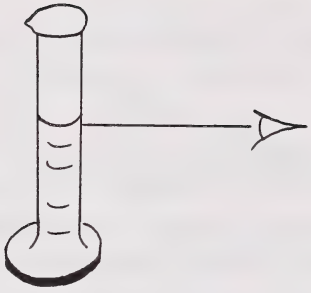


Figure 1: Reading the level of the liquid.

When transferring solids from one place to another, always work with clean instruments, so you don't have any reaction with unwanted particles. Also, if you take material from a stock bottle, never transfer solids back to it in order to avoid contamination. Furthermore, the cap of a bottle should always be placed on the desk with the top down to avoid contamination. Try to estimate fairly closely the amount of solid you will need from a bottle.

Should you ever have to filter a precipitate, take some round filter paper; fold it in half; and then fold it in half again. The round paper should then look like a quarter of a pie. Then, open one end and put it into a funnel, and pour your solution through it. The liquid portion should go through, and the solid portion should be caught by the filter paper.

If a solid is completely dissolved in a liquid, rather than suspended in a liquid, this solid can be recovered by evaporating the liquid.

Many things can be done with glass, however, just some things will be mentioned here. If you break a glass tube, and then wish to get rid of the sharp edges, you can fire-polish it. When you fire-polish glass, you melt the edges and make them smooth.

When chemicals are weighed on a balance, do not put the chemical on the balance directly, but put them on a piece of plastic or lightweight paper. Also, don't weigh hot objects, and keep the balance dry.

Exercise 1

1. When weighing chemicals on a balance the metal weighing pan should be protected by _____.

2. The surface of liquids when viewed in a glass cylinder is always curved. This lens-shaped surface is called a(n) _____.

3. For a most accurate reading of a liquid in a graduated cylinder the meniscus must be viewed along _____

_____

4. In transferring solids from reagent bottles list three general rules to be observed in order to avoid contamination of reagents.

(a) _____

(b) _____

(c) _____

5. When a solid dissolves in a liquid (salt dissolved in water) the solid can be recovered by _____.

6. What is meant by "fire-polishing" of glassware?

_____

_____

_____

7. List a practical purpose of fire-polishing.

_____

C. Experiment 3 - Qualitative and Quantitative Observations

Observations may be of two types: qualitative and quantitative. A qualitative observation occurs when a description is specified as to kind, but not to amount. A quantitative observation occurs when you specify what happens in terms of numbers. For example, saying it rainedhard yesterday is a qualitative observation. However, saying that 4 cm of rain fell between 3:30 and 5:30 is a quantitative observation.

In this experiment, you will gain practice in making good observations. Acquire the necessary materials, and then follow the procedure. As you do so, fill in the table in the appropriate place, indicating all of your observations.

Purpose: To make and record meaningful observations while watching materials change.

Materials: 3 sugar cubes
3 glasses
watch or clock
hot and cold tap water

Procedure:

Fill one glass with extremely hot tap water, and fill the other two glasses with cold tap water. Completely crush one of your sugar cubes. Then add a whole sugar cube to the hot water, add a whole cube to one glass of cold water, and add all of the crushed cube to the last glass of cold water. Watch what happens and record the rate of which various changes take place, as you fill in the table in the appropriate places. All sugar should be placed in the water very rapidly so that accurate comparisons can be made.

Exercise 2

Tabulate your observations in the table below.

| | Cold Water | Hot Water | Crushed-Cold |
|---|---|---|---|
| Time: | | | |
| 20 s (seconds) | | | |
| 2 min | | | |
| 5 min | | | |
| 10 min | | | |
| 30 min | | | |
| 60 min | | | |

Exercise 3

1. From your observations of Experiment 3, please write the word Qualitative or Quantitative after each of the statements given below.

   (a) In hot water sugar dissolves faster than in cold water.

   _____

   (b) In 10 min exactly one-half of the sugar cube dissolves.

   _____

   (c) All cubes started dissolving after a short period of time.

   _____

   (d) A sugar cube, before dissolving, is 1.5 cm x 1.5 cm.

   _____

   (e) Bubbles form as sugar cubes dissolve. _____

   (f) In 50 min the sugar cube in hot water dissolves completely.

   _____

2. Fill in the blanks with appropriate terms or sentences.

   (a) Did you observe any bubbles rising from the whole sugar cube?

   _____

   (b) Did you observe any bubbles rising from the crushed sugar cube?

   _____

   (c) Give a possible interpretation for the occurrence of bubbles when sugar cubes dissolve.

   _____

   _____

   _____

D. Graphs

Graphs are a convenient way to show data at a glance. With tables or words, any changes are not nearly as apparent as when the changes are plotted on a graph. Graphs generally have two axis. One variable is plotted along the horizontal axis, and the other variable is plotted along a vertical axis. At a glance, it can be seen how a change in one variable affects the next item. The data that you collect will give you a number of points which should then be smoothly connected. Your resulting straight line or curve should go between all points, since this will tend to cancel out experimental errors at any point. Incidentally, this type of graph is a line graph, but you can also have a bar graph, or a graph in a circle, or other types of graphs.

When making a line graph, you must always have a title. The horizontal axis should always be labelled with the independent variable. This is the variable that is determined in advance, and which doesn't change. For example, if you want to see how much you grow every year, the years would be plotted along the horizontal axis. The dependent variable, or the one which changes, such as your growth rate, is plotted along the vertical axis. In the experiment below, a solid was heated and a liquid was cooled. The temperature changes were recorded and plotted on a graph.

E. Experiment 4 - Melting and Freezing of Solids

A phase change occurs when a solid is heated and melts to a liquid. When the liquid is re-cooled, it solidifies at the same temperature at which the solid melts.

Consider the heating and cooling curves in the time-temperature relationship for ice shown in a sample experiment below.

Record:

Ice from a refrigerator was heated over a limited range by a constant source of heat. The temperature of the sample confined in a test tube was noted every one-half minute.

Similarly water in a test tube was cooled and frozen and the temperature noted every one-half minute.

Data:

Phase Change: ice ⟶ water

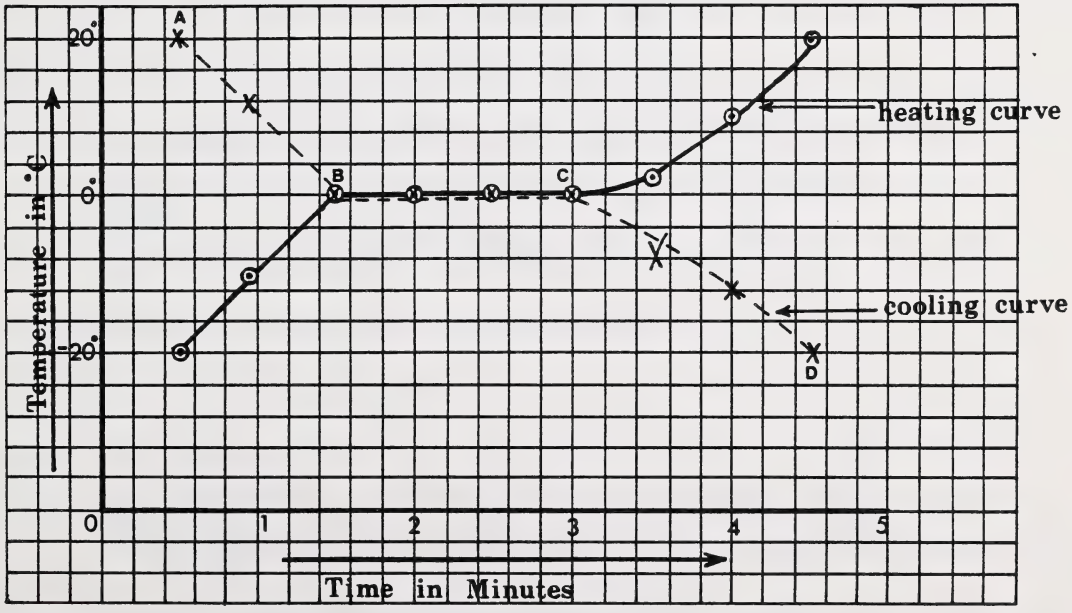
| Time Minutes | Heating data °C | Phase |
|---|---|---|
| 0.5 | -20° | solid |
| 1.0 | -10° | solid |
| 1.5 | 0° | |
| 2.0 | 0° | phase |
| 2.5 | 0° | change |
| 3.0 | 0° | |
| 3.5 | 2° | liquid |
| 4.0 | 10° | liquid |
| 4.5 | 20° | liquid |

Phase Change: water ⟶ ice

| Time Minutes | Cooling data °C | Phase |
|---|---|---|
| 0.5 | 20° | liquid |
| 1.0 | 12° | liquid |
| 1.5 | 0° | |
| 2.0 | 0° | phase |
| 2.5 | 0° | change |
| 3.0 | 0° | |
| 3.5 | -8° | solid |
| 4.0 | -11° | solid |
| 4.5 | -20° | solid |

From the data a graph was obtained by plotting the time against the temperature.

Cooling and Heating Curves for Ice



Interpretation of curves:

AB on the cooling curve shows the time taken for the water to cool from 20°C to 0°C. At 0°C, the freezing point, ice starts to appear. With further cooling, freezing continues until all the water is frozen. While freezing is taking place, the temperature remains constant as shown by BC, the plateau on the graph. Continued cooling causes a drop in temperature of the ice as shown by CD on the graph.

A similar curve is shown for the heating of ice. Note that there is also a plateau, or a line of constant temperature, on the heating curve. This represents the time required for all the ice at 0°C to change to water at 0°C.

Increasing the amount of ice used in the experiment would not change the melting temperature but the slope of the heating curve would be more gradual and the plateau longer. Applying more heat per minute would cause a more rapid rise in temperature and would shorten the length of the plateau.

An observation is something that you can see or hear or feel etc. When "observing", you are not thinking about why something happens the way it does. An interpretation is when you try to give a reason for your observations.

Exercise 4: Observation, Interpretation, and Conclusion

1. Describe the graph (note changes in its shape).

_____

_____

_____

_____

2. Give your interpretation of the shape of the heating and cooling curves.

_____

_____

_____

_____

3. Based on the data, what is the melting temperature of ice? _____

4. Based on the data, what is the freezing temperature of water? _____

5. What effect would increasing the amount of ice have on the shape of the heating or cooling curves?

_____

_____

_____

6. (a) Does a flat region (plateau) occur in the graph? _____

(b) If a plateau does occur, give a possible interpretation for this occurence.

_____

_____

7. Conclusions:

Is the substance a pure substance or a mixture?

_____

Explain.

_____

_____

_____

## Exercise 5

Which of the following are observations and which are interpretations.

1. (a) The plateau on the cooling curve of a liquid indicates a phase change. _____

(b) Sugar dissolves faster in hot water. _____

(c) Bubbles rise from dissolving sugar cubes. _____

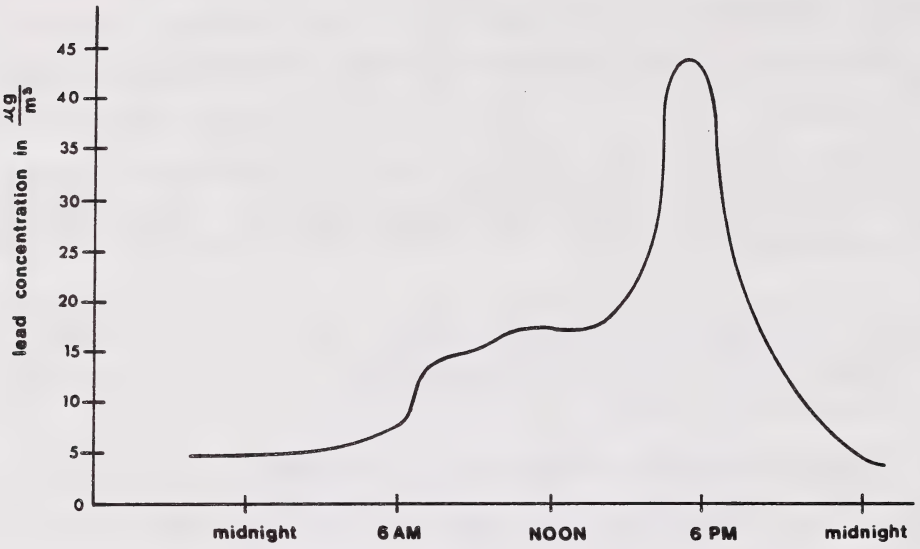
(d) Heat and light are produced by a burning candle.

_____

(e) The bubbles that rise from the dissolving sugar cubes are air trapped in the cubes, and not gas due to reaction, since no bubbles formed with crushed sugar.

_____

2. Following is a graph of lead concentration in the air of a large city, in the heart of the city, during a 24 h period.




Classify the following three observations as either qualitative or quantitative.

(a) The concentration increases as more cars are on the road. _____

(b) During the afternoon rush hour period, the concentration reaches 45 $\mu g/m^3$. _____

(c) There is always some lead in the air, regardless what time it is. _____

(d) Is the statement in (a) an observation or an interpretation? _____

(e) What is the concentration of lead at 3:00 p.m.? _____

## F. Tests for Some Common Gases

Various tests can be used to identify different substances. For example, if you have certain metal ions in a solution, certain reagents can be added. These chemicals will react in a certain way with some metal ions, and they will react differently, or possibly not at all with other ions. Depending on the reactions you get, and on the colors of the various precipitates you can often positively identify a given ion.

Different tests can also be used to identify different gases, since different gases have different properties. For example, hydrogen itself will burn, but it will not support burning. So if you have a reaction in which gases are bubbling off, the gases can then be collected in an inverted test tube. A burning splint can be put into the test tube, and if you hear a barking sound, you know the gas is hydrogen. This is because the burning splint ignites the hydrogen, causing a minor explosion, hence the barking noise.

Carbon dioxide will turn limewater milky. Limewater is $Ca(OH)_2$ dissolved in water. The carbon dioxide would react with $Ca(OH)_2$ in the following reaction:

$$CO_2 + Ca(OH)_2 \longrightarrow CaCO_3 + H_2O$$

The milkiness is due to the $CaCO_3$ forming a precipitate in water.

As you all know, oxygen makes everything burn faster, which is why no smoking is allowed in places where oxygen is being used. To see if oxygen has been collected in a test tube, light a splint, and then blow it out so it is just glowing. Insert it into the test tube. If the splint bursts into flames you know you have oxygen.

Exercise 6

Briefly describe the test for the following gases:

1. Hydrogen: _____

_____

_____

_____

2. Oxygen: _____

_____

_____

_____

3. Carbon dioxide: _____

_____

_____

_____

Note: You may have experienced some of these tests in former science courses.

END OF LESSON J